





THE BOYS' OWN BOOK OF SCIENCE

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PREFACE

THE BOYS' OWN BOOK OF SCIENCE has been written for that host of boys everywhere who wish to experiment at home. It is not a book about science. It is a practical guide to real worth-while experimental work. Under its direction the amateur will come to feel that he is not only acquiring knowledge, but that he is becoming a real scientist. The book has been written out of an experience of twenty years in teaching and in directing the home activities of boys in this sort of work. As an inspiration to their efforts, there have been included sketches of a number of world-famous scientists who started as home laboratory workers.

It is not intended that every boy will do all of the experiments in this book. There is an abundance of material from which to select. Each boy may choose for himself that which most appeals to him. It will afford a browsing ground for a long period of time. Gradually real equipment may be accumulated and a genuine laboratory established. The laboratory is bound to grow with increased skill and a growing knowledge of fundamental processes. This book is not intended merely to entertain, although there will be entertainment in abundance. Its purpose is to direct

the latent energies of boys who wish to explore for themselves the fields of scientific discovery. Properties of matter, the nature of chemical and physical forces, and methods of preparing elements and compounds, as well as processes of analysis, have an irresistible appeal for all boys. This book has been written to satisfy that very legitimate interest.

As a source book of information and experimental hints for teachers, this book may well be useful. Although not written for them, it contains much that will be of value, particularly to the beginning teachers of science.

FLOYD L. DARROW.

Brooklyn, N. Y.

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CHAPTER I

THE HOME LABORATORY

THE location of the home laboratory is one of the most difficult problems that confronts the amateur experimenter. Most homes are not planned with any provision for his convenience and comfort. And yet, if at all possible, he should have a room—a small one will do—undisturbed by the rest of the family, where he may have perfect freedom to do real laboratory work. An upper room is preferable, and it should be as light and airy as possible. Of course a kitchen may be used as a laboratory, but apparatus and chemicals cannot be left standing around in it. And it is always a nuisance to have to get everything out and put it all away every time you wish to experiment. If possible, by all means have a permanent place to work.

Now as to the equipment of your laboratory, that is a most important matter too. Across one side of your room build a bench about 30 inches high and 2

feet wide. The top should be of hard wood and planed smooth. Paint it black and if possible cover it with slabs of plate glass. The glass will not be affected by any chemicals that you may use, and it can be easily kept clean. At one end of the bench you ought to have a sink with running water and waste pipe. Of course that is ideal. Many of you cannot have such luxurious quarters, and in that case you must bring water into your laboratory and carry it out each time that you experiment. At the bench should be a hose connection for illuminating gas. If your room is lighted with gas, this can be easily arranged. If you use electricity but have gas in the house, a plumber could probably run a pipe to your room without much trouble. Of course if gas is out of the question, you will have to substitute an alcohol lamp. And there are a number of good alcohol lamps on the market now which compare favorably in heat-giving power with a Bunsen burner. At the back and running the whole length of your bench build some shelves to hold your apparatus and chemicals. These do not need to be very deep. Six inches will be sufficient. As your laboratory grows and you accumulate larger pieces of apparatus, a cupboard may be built in another part of the room.

The selection of apparatus and chemicals will require better judgment than any other feature of your equipment. It is usually not best to buy a regular chemical outfit. At the start such an outfit will be more or less of a mystery to you. Frequently the apparatus is not

of the highest quality and the quantities of chemicals are small. The most economical and satisfactory method is to select a certain number of experiments from the following chapters which you wish to do. Note the pieces of apparatus and materials which you will need and then order them of a reliable firm. You will find that many of these materials are very common. Some of them you will find in your own household. Others may be had from the druggist or the grocer. A large granite basin and wide-mouth bottles will serve for the collection of gases. Apparatus must be ordered from some chemical company.

But if you build your laboratory in this way, it will grow with your knowledge. You will know the purpose of every piece of apparatus and every chemical. When you are through, your laboratory will represent a personal creation. Your experience will have genuine educational value and not be a mere entertaining pastime of which you will soon grow weary.

To aid you in the selection of apparatus I have arranged two cuts that show many of the principal pieces which sooner or later you will need. If at any time you do not know what is meant by a certain piece of apparatus, refer at once to these illustrations and their names in the footnotes.

Of course you will need a Bunsen burner or an alcohol lamp, a ring-stand, a wire gauze with asbestos center, some test tubes, a few rubber stoppers of various sizes, several lengths of glass tubing about five-sixteenths of an inch outside diameter, four feet



FIGURE 1.
LABORATORY APPARATUS.

1, nest of beakers; 2, Bunsen burner; 3, hornpan balance; 4, retort; 5, set of weights; 6, test-tube holder; 7, burette; 8, pipette; 9, thistle tube; 10, tripod; 11, Erlenmeyer flask; 12, gas generator; 13, water bath; 14, blowpipe; 15, wash bottle.



FIGURE 2.

LABORATORY APPARATUS.

16, test-tube rack; 17, pneumatic trough; 18, steel forceps; 19, ring-stand; 20, crucible and lid; 21, funnel; 22, iron deflagrating spoon; 23, pipemeter triangle; 24, test-tube brush; 25, mortar and pestle; 26, clamp; 27, alcohol lamp; 28, glass cylinder; 29, porcelain evaporating dish; 30, pinch-clamp, 31, flask.

of rubber tubing to fit, a flask of 500 cubic centimeters' capacity, an Erlenmeyer flask half as large, a small evaporating dish, a mortar and pestle, a pair of steel forceps, and a porcelain crucible and lid. Other needs will present themselves as you proceed with the work.

All volumes of liquids will be measured in cubic centimeters and weights of chemicals will be in grams, since the metric system is everywhere used in scientific work. You will need a graduate of about 25 cubic centimeters' capacity and a hornpan balance and weights. These need not be expensive. The weights should go from 1 gram to 50 with additional decimal weights running from 0.01 to 0.5 of a gram. In all future directions I shall use the abbreviations c.c. for cubic centimeters and g. for grams.

Another quite important item of your equipment will be the matter of electricity. For some of the work, I will direct you later how to make simple cells that will be perfectly satisfactory. But if you are to do any of the work that I describe with electric furnaces, another source of supply will be essential. If you have electricity in your house, the problem may be easily solved. Wires may be run to a couple of binding posts on your bench and you will have available at all times 110 volts. For electric heat, either direct or alternating current will serve your purpose. If the voltage is to be cut down by a resistance and the current used for most other work in place of cells, direct will be required. You must remember always that in using a 110-volt source of current a considerable

resistance must be placed in series with your apparatus, if you are not to blow the fuses or possibly ruin your apparatus. But I will give you directions about resistances when the experiments are described. If it is not convenient to run wires to your bench, a connection may be made to a lamp socket. Between the binding posts on your bench or the wires from the lamp socket and your apparatus always place a switch. Then if trouble arises you may easily break the circuit.

In starting this work make up your mind that you are going to become a real home laboratory worker. Do not let your chief purpose be amusement only, although genuine enjoyment you will get in abundance. Have in your laboratory a shelf of the latest texts of elementary chemistry and study the subject as you proceed. Add to your library whenever possible. You will be surprised and delighted to see how rapidly your working knowledge of chemistry will expand. And it may quite possibly be that this amateur work will prove to be a stepping-stone to your future career. The director of one of the largest technical schools in New York said in a letter to me the other day that the demand for men with laboratory training far exceeds the supply. You may qualify to meet this demand, and your initial training may be had in your own laboratory.

CHAPTER 2

LABORATORY MANIPULATION

THERE is much to learn regarding the care of apparatus and its use in the laboratory. Gradually you will acquire the technique of laboratory practice. At the start a few simple directions will be of value.

The Burner.—Examine your Bunsen burner and note how it is made. Observe that there are holes at the bottom of the tube for the admission of air, and an air adjustment. In lighting the burner turn the gas on full and hold the match two or three inches above the top of the tube. If these precautions are not observed, the flame will unusually “strike back” and burn at the base. Whenever this happens, turn off the gas immediately and relight it. If you wish to obtain a small flame, with the air adjustment cut off some of the air and at the same time turn down the gas. Note that when the air-holes are completely closed, the flame is luminous. Such a flame deposits soot and is seldom used. Always use the clear blue flame.

Glass Cutting.—Very often you will have occasion to cut glass tubing. To do so make a slight groove with a triangular file at the point where you wish to cut



Bending glass tubing with the "fish-tail" flame.

it. Then grasping the tube with both hands, the thumbs meeting opposite to the groove, bend the ends of the tube toward you. This should give an even break.

Before using glass tubing, fire polish the ends. To do this hold each end of the tube in the flame for a few moments rotating it, as you do so. The glass will soften and leave a smooth edge in place of the sharp ragged one.

Glass Bending.—Another very important operation is the bending of glass tubing. This will be necessary in making angle-tubes for rubber stoppers and delivery tubes for collecting gases.

In bending glass place a “fish-tail” top on your burner so as to obtain a broad flat flame. Then holding the tubing lengthwise of the flame rotate it with your fingers in order to heat it evenly on all sides. In a few moments the glass will soften and, when the flame is colored an intense yellow, you may remove the tubing and bend it quickly to the desired shape. The bend should always be broad and round, instead of sharp and angular. Then it will not easily break:

Stoppers.—For most purposes rubber stoppers are best. They come in sizes numbered from 0-0 to 10 and even higher. For your work the first six sizes will probably be sufficient and you should have two or three of each. It is always best to get two-hole stoppers, except in the first three sizes. Then if you need a one-hole stopper you may easily close one hole with a glass plug. To make a plug hold with the tongs an inch

length of glass tubing vertically in the flame until the opening is closed.

In passing a piece of glass tubing through a stopper do not force it. That will usually break the tubing. Moisten both tubing and stopper and the stopper will slip on easily.

Corks of various sizes may probably be obtained from your druggist. For making holes in them you will need a set of cork-borers which you will readily see how to use. To bore a hole in a rubber stopper wet the boring tube in dilute sodium hydroxide solution.

Heating Glassware and Porcelain.—A test tube may be held directly in the flame, but beakers and flasks must be protected by a wire gauze. It is best to have one or two 4-inch squares of gauze with an asbestos center. In heating any glass apparatus containing a liquid, never allow the flame to extend above the level of the liquid. Vessels like bottles and battery jars having thick walls can never be heated. Porcelainware may be heated to a high temperature, but the heat should be applied gradually. Crucibles may be put directly in the flame but evaporating dishes should be protected by a wire gauze.

General Hints.—Before you actually begin to perform an experiment have everything in readiness.

Clamp all glass apparatus loosely, especially if it is to be heated.

Do not force stoppers into the necks of thin glass flasks or test tubes.

Always be sure that the lower end of a thistle tube dips beneath the liquid in the flask.

In removing a stopper from a reagent bottle do not lay it down. With the palm of the hand turned upward, remove the stopper by pressing it between the lower joints of the first and second fingers. Then grasp the bottle between the thumb and lower fingers. After pouring out the reagent remove the drop clinging to the lip of the bottle by touching it to the receiving vessel and replace the stopper.

Folding Filter Paper.—You will frequently have occasion to filter liquids. Fold the circular piece of filter paper in half. Then fold it in half again. When opened up this will exactly fit the angle of your funnel. Press the filter against the sides of the funnel, and moisten it with a few drops of water to hold it in place. Filtering is a slow process, but a properly fitted filter will greatly hasten the operation.

Quantities of Material.—Do not adopt the plan that “if a little is good more is better.” In making tests, very small quantities give much more decisive results than large ones. Wherever the amounts are not specified, use small ones.

CHAPTER 3

THE ALCHEMIST

THE original home-laboratory experimenter was the alchemist. Possibly you have never heard of him. He was a curious type of scientist, but there were many of them scattered throughout Europe all during the Middle Ages and even down to quite modern times. They worked in dark caves and hidden places, keeping their discoveries secret from other men.

There was a mythical story abroad during those early centuries about a wonderful substance known as the "Philosopher's Stone." Whoever should be able to find or make this magic stone would be in possession of everything that men might desire. At its touch the baser metals would at once turn into gold and its mere presence would give to the fortunate possessor perfect health and perpetual youth. In the ignorance and superstition of those days no tale seemed too fantastic for belief. Indeed, there seemed much evidence in support of this idea of changing one metal into another. Iron rusted, and, as they thought, a new substance formed. If a rod of iron were thrust into a solution of blue vitriol, the iron disappeared and metallic copper took its place. What more natural

than to suppose that iron had been transformed into copper? Were not gold and silver actually found in lead ore, and did not this indicate that these precious metals were transmuted lead? And, too, it was believed that worked-out mines after long periods of time exhibited new "crops" of gold and silver. Could they not learn the secret of this change and be able to control it? For centuries men pursued this will-o'-the-wisp. Strange, it seems now, doesn't it?

Many was the impostor who rose to prey upon the credulity of king and prince and obtain ill-gotten gain under the false pretense of having found this magic substance. An old alchemist was one day giving a dinner to some friends when, with much unconcern, he related an incident from his personal experience that had occurred eight hundred years before. Noting the looks of amazement upon the faces of his guests, he turned to his servant and said, "John, is not that true?" And the faithful servant replied, "You forget, master, that I have been in your service but five hundred years."

But although the philosopher's stone was never found, the search for it resulted in much gain for chemistry. New substances were found or made. These workers became expert in laboratory methods. Many of the new drugs were useful as medicines, and gradually there grew out of alchemy the art of the apothecary. In fact this search for the healing properties of new compounds in time became the chief purpose of chemistry. But in their eagerness to discover

the curative effects of some new-found drug these crude practitioners hastened to administer it to some patient instead of first trying it out on a cat or dog. Indeed, it was in this way that the fatal properties of our violent poisons were made known.

But the real chemist in time took the place of this misguided seeker after the impossible, and soon you in your home laboratory will be in possession of more real chemical knowledge than all the alchemists ever knew.

CHAPTER 4

CHEMICAL MAGIC

I AM sure the users of this book of home laboratory experiments are not going to be much different from the large number of other boys whom I have been directing in this kind of work for a number of years. It is the mysterious and wonderful things in chemistry and physics that always appeal to beginners. So I am going to start you off with some experiments on chemical magic. Gradually we shall come to more serious things and work of real chemical importance. The experiments in this first chapter will give you an abundance of material, not only for your own amusement but for parlor entertainments.

The National Colors.—Let us begin by being patriotic. Place on the table before you a glass pitcher and three tumblers. Fill the pitcher with water into which you have stirred a few c.c. of a solution of ferric chloride. To all appearances the water will be nearly colorless. In the first tumbler place two or three drops of a solution of ammonium sulfocyanate. Leave the second tumbler with nothing in it. In the third, place two or three drops of potassium ferrocyanide solution. You should have these preliminaries arranged in advance. Then announce to your audience that you will

pour the national colors from a pitcher of pure water. Upon filling the tumblers, the first will be a deep red, the second colorless, and the third a bright blue. Like all trick experiments this will be perfectly mystifying to the uninitiated but quite clear to you. The iron compound used, when present even in small quantities, gives with these indicators the characteristic colors observed.

“Freezing” Water by Magic.—Into a small flask or breaker put 50 g. of ordinary photographer’s hypo and 10 c. c. of water. (A small hornpan balance and graduate will enable you to get these quantities without difficulty). Heat this mixture over your alcohol or Bunsen burner until the hypo just dissolves. Then let it cool carefully without disturbance. A clear liquid solution will result. Have at hand a tiny crystal of hypo about as big as a pinhead. Holding the flask before your audience, state that in the twinkling of an eye you will freeze this water to solid ice. Dropping the crystal into the solution, give the flask a quick shake and immediately its contents will change to a mass of solid crystals. At the same time the flask will become decidedly warm giving experimental evidence of the fact that when water freezes or a solid crystallizes heat is given off. You have prepared in this experiment what we call a *supersaturated solution*, and when the crystal is added the water recombines chemically with the hypo to form the original crystals again. But to your observers it appears like the actual freezing of water.



Igniting the wick of an alcohol lamp with a "glass rod."

Igniting Gasoline with a Glass Rod.—Place on your demonstration table a small pane of window glass, a beaker, and a glass rod. In the bottom of the beaker put a half teaspoonful of potassium permanganate crystals and moisten them with a few drops of water. Pour on the window pane two tablespoonfuls of gasoline. Add two or three cubic centimeters of the strongest sulfuric acid to the contents of the beaker. Immediately a vigorous chemical action will be set up. Then dip the glass rod into the sputtering mixture and touch it to the gasoline. With a flash the gasoline will take fire and blaze up.

If the rod is again dipped into the mixture and touched to the wick of an alcohol lamp, the wick will light. The chemicals in your beaker have produced ozone, a very concentrated form of oxygen, and this has oxidized the gasoline and alcohol rapidly enough to bring them to their kindling temperatures.

Some Mystifying Color Experiments.—Arrange three rows of tumblers on your table and a pitcher of water containing one or two tablespoonfuls of strong household ammonia. In the bottom of each tumbler in the first row place two drops of phenolphthalein indicator. This indicator is prepared by dissolving a tiny pinch of phenolphthalein powder in a little denatured alcohol. To each tumbler in the second row add a few drops of strong sulfuric acid. Two or three will probably be enough. To the tumblers in the third row add a few drops of a strong solution of sodium hydroxide, or lye.

Begin to fill the first row of tumblers and behold from your pitcher of "water" will apparently come forth "wine." Then, stating that you wish to get rid of this dangerous evidence, pour the contents of the first row of tumblers into the second. Immediately the wine changes back to water. And, upon pouring the water into the third row of tumblers, the wine reappears.

Just exactly what has happened will become clear when we have had some later work on acids and bases.

Freezing a Tumbler to a Block of Wood.—Upon a smooth block of wood place a few drops of water and set in it a thin-glass tumbler. Fill the tumbler an eighth full of water and stir into it ammonium nitrate powder in considerable quantity. Presently moisture will condense upon the outside of the tumbler, then frost will appear, and finally the tumbler will freeze to the block so that you can pick it up and the block will not fall. The solution of ammonium nitrate produces great "cold."

This same stunt may be done by blowing through a glass tube into ether or carbon disulfide placed in the tumbler. The evaporation of the liquid also produces cold.

Obtaining Water from Solid Crystals.—Place in a dry test tube a few crystals of native gypsum, that is, gypsum just as it comes from the mine. Heat the test tube holding it so that it will slant gently downward. Almost immediately steam will appear and moisture will condense on the cold walls of the tube. Presently

drops of water will drip from the mouth of the tube. And there is a most remarkable fact about this water. It has been locked up in these crystals for millions of years. Any gypsum that is now quarried was formed in the earth's crust ages ago, and the imprisoned water fell as rain in a far distant past. And yet the water is the same in composition as water produced now. It

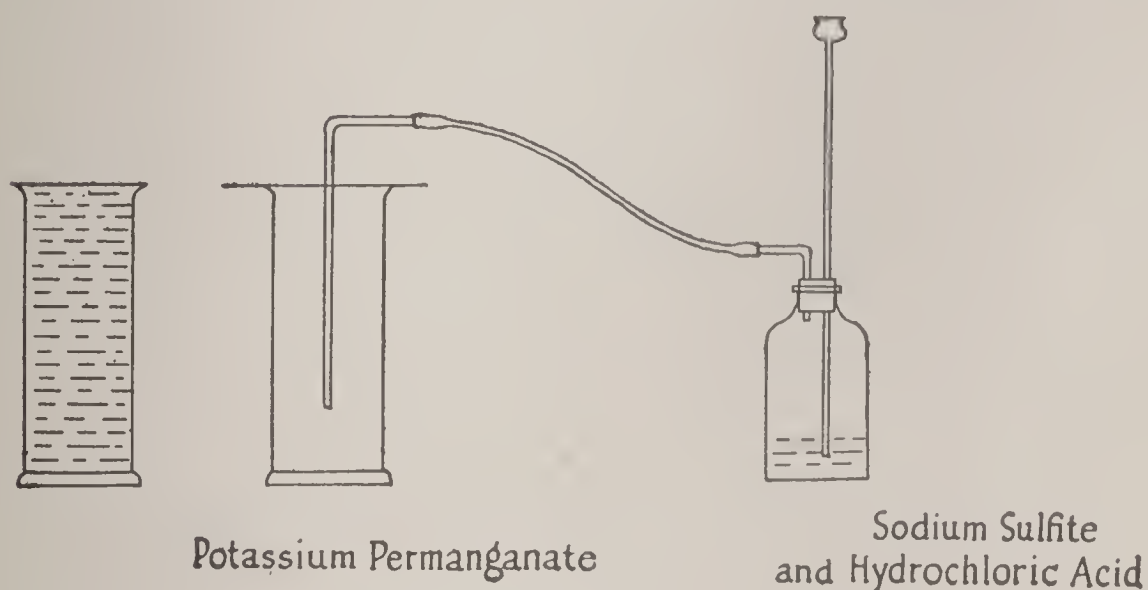


FIGURE 3.

Apparatus for preparing sulfur dioxide and converting "wine" into "water."
Potassium permanganate is in cylinder at left.

is perfectly pure and, could you get enough of it, fit to drink. Upon examining the residue in the tube you will find that the gypsum has lost its crystal form and may be crumpled to a powder.

The same experiment may be repeated with crystals of blue vitriol and photographer's hypo. But in these cases the salts have been artificially prepared and the water is not of ancient origin.

Changing "Wine" into Water.—There is another color experiment more baffling than any of the pre-

ceding. In a clear-glass pitcher prepare a purple solution, resembling wine, by pouring water onto a few crystals of potassium permanganate. Then set up a sulfur dioxide generator as shown in figure 3. The generator may be a wide-mouth bottle to which you have fitted a thistle tube and delivery tube. In the generator place sodium sulfite and having arranged for the collection of the gas, as indicated, pour through the thistle tube a little hydrochloric acid. Immediately a vigorous action will begin and heavy sulfur dioxide gas will pass through the delivery tube and fill the receiving cylinder or bottle. You can tell when the cylinder is full by the abundance of sharp choking fumes that come from about its mouth. In filling a cylinder with gas in this way, always pass the bent delivery tube through a perforated cardboard cover.

Now remove the delivery tube and cover the cylinder with a glass plate. Have the gas generated in advance of your audience. Since sulfur dioxide is a colorless gas, the cylinder will appear to be perfectly empty. Then announce that you are to accomplish the magic transformation of wine into water. Upon pouring the purple potassium permanganate solution into the sulfur dioxide cylinder, the color will disappear just as fast as you can pour it, and apparently water will have formed. The sulfur dioxide really reduces the permanganate to a colorless compound.

The Magic Wand.—You will find it easy to produce a color change which will appear only after an



Generating sulfur dioxide, preparatory to changing "wine" into "water."

interval of time and at the precise moment that you wave your magic wand.

Prepare a solution of sodium iodate by dissolving one gram of the substance in a pint of water. Stir into this a few c. c. of thin starch paste made by pouring boiling water on to a pinch of starch and allowing it to cool. Fill a large tumbler half full of the mixture.

Then using the same apparatus as that of the previous experiment, prepare a solution of sulfur dioxide by bubbling the gas for a few moments through a half pint of water.

Now pour the sulfur dioxide solution into that of the iodate and mix thoroughly. Nothing seems to happen at first, but in from one to five minutes, depending upon the strengths of the solutions, a deep purple color will suddenly spread throughout every portion of the mixture. By timing this interval you will be able to determine just exactly how many minutes and seconds will elapse between mixing the solutions and the appearance of the color. Then at the psychological moment wave your wand and behold the color.

Changing "Air" to Brown Fumes.—Using the sulfur dioxide generator, arrange your apparatus as shown in Figure 4. In the generator place a few copper rivets. For the pneumatic trough, over which to collect the gas, a granite basin containing an inch-depth of water may be substituted. Fill a wide-mouth bottle with water and covering it with a glass

plate invert it in the basin, leaving the bottle filled with water.

Now nearly cover the copper rivets with water and pour through the thistle tube strong nitric acid. Soon the generator will begin to bubble and fill with brown fumes. Then placing the delivery tube beneath the

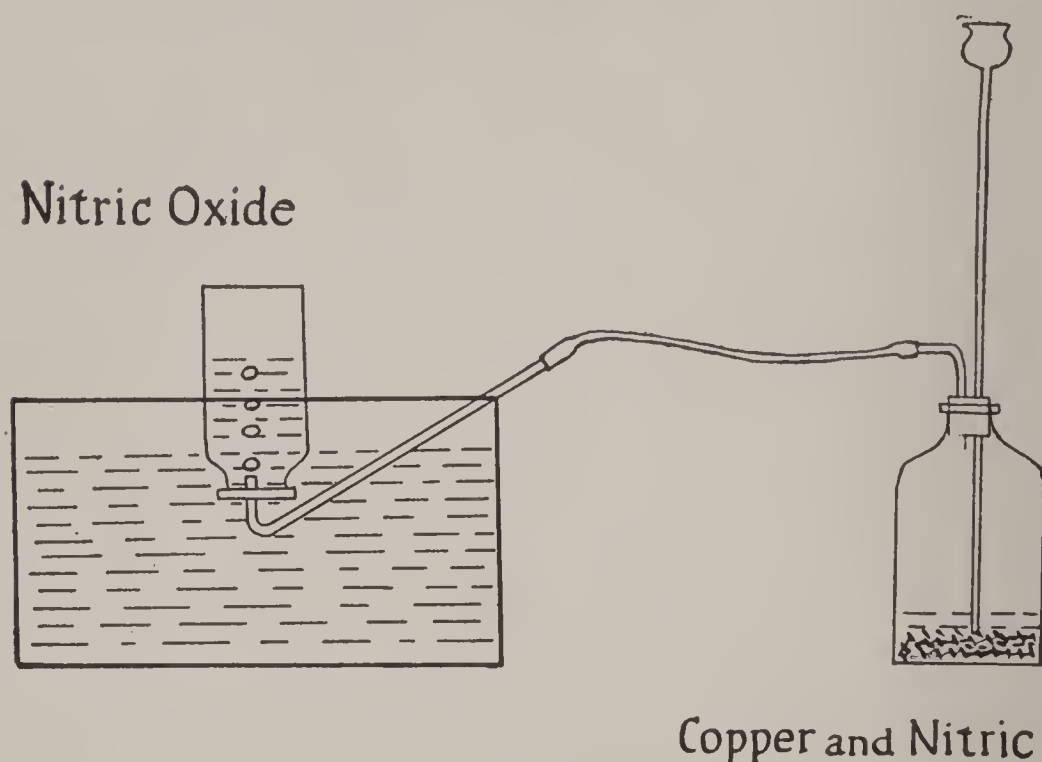


FIGURE 4.

Preparation of nitric oxide for changing "air" to brown fumes.

slightly upturned mouth of the bottle, fill the bottle with gas, which will rise through the water and displace it. If you like, several bottles may be filled with the gas. To remove them place glass plates beneath their mouths and invert them upon the table.

The gas in the bottles is colorless, but, upon removing the glass plates so that the air may act upon it,

dense brown fumes appear. It seems like changing the color of air.

If a little water is shaken in the bottle, the fumes dissolve giving a colorless solution, which turns red upon adding a solution of blue litmus.

Something from Nothing.—Dissolve three or four tablespoonfuls of granulated sugar in as little water as possible. Place the solution in a tall tumbler or jar. Then pour into it an equal quantity of the strongest sulfuric acid. Immediately the mixture turns black, begins to steam, and presently it boils up and fills the tumbler to overflowing with a thick black mass of frothy carbon, while the odor of burnt sugar fills the air. It will be amazing to see the large mass of solid material that will come from so small a quantity of sugar.

The sugar consists of carbon together with hydrogen and oxygen in proportion to form water. Sulfuric acid has a very great affinity for water and removes the hydrogen and oxygen leaving the black carbon.

Queer Forms of Sulfur.—Ordinary sulfur, or brimstone, is easily obtained. Fill a test tube two-thirds full of broken pieces and heat them very slowly in the flame, taking the tube out frequently to prevent a too rapid heating. Gradually the sulfur will melt into a clear straw colored liquid. Then upon further heating this liquid will change to a thick cherry-red substance which will not flow from the test tube even when inverted. It looks like jelly and good enough to eat. When heated again, the color becomes nearly black,

and the sulfur changes to a thin boiling liquid. Have at hand a jar of water and quickly pour the boiling sulfur in a thin stream into the water. Probably the sulfur vapor will take fire, giving an example of spontaneous combustion, but don't be alarmed at that. Remove the sulfur from the water and you will find a rubbery mass, which will stretch like elastic. Upon standing for twenty-four hours it will change back to the original brittle form.

Sympathetic Ink.—Dissolve a pinch of cobalt chloride in a little water and write your name with it upon white paper. When dry, nothing will be visible. But, if the paper is held high above the flame of an alcohol or Bunsen burner, the letters will appear in blue lines. As it cools, the letters will gradually fade.

You may prepare in advance some mysterious message and state to your audience that you have a spirit communication which will appear only under the influence of fire. Show them that the magic paper has no writing upon it. Then go through certain incantations at the same time holding the paper above the flame. Presently the message will be revealed.

Magic Writing Papers.—Prepare three mixtures, each of which will contain powdered ferric ammonium sulfate, and in addition in number one an equal quantity of powdered potassium ferrocyanide in number two tannic acid, and in number three ammonium sulfocyanate.

Keeping all materials perfectly free from moisture, rub one piece of white writing paper over thoroughly



Filling a flask with ammonia for a fountain in a vacuum.

with mixture number one, another with mixture number two, and a third with mixture number three. Shake off any loose powder and then write with water upon each paper. The letters will appear blue in the first, black in the second, and red in the third.

Changing one Metal into another.—Prepare a solution of blue vitriol and fill a cylinder or tall narrow bottle with it. Then thrust into the solution a strip of sheet zinc. Immediately the zinc will become discolored and gradually it will disappear. At the same time the blue color disappears from the solution and a deposit of red copper will accumulate in the cylinder. The zinc has gone into solution and displaced the copper. The alchemists regarded this as the actual transformation of one metal into another.

Producing a Magic Fountain.—Arrange apparatus as shown in the illustration. Fill the flask over the Bunsen burner a third full of the strongest household ammonia. In the large glass jar beneath the inverted flask place water containing litmus solution and two or three drops of sulfuric acid. This will color the water red. Have a one-hole rubber stopper carrying a glass tube extending up into the flask arranged as shown and dipping into the liquid in the jar below. Then heat the ammonia and drive the gas through the delivery tube into the inverted flask. Since the gas is light, it will rise and displace the air. When the flask is full, which may be told from the strong odor of ammonia about the mouth of the flask, inject into the flask a little water from a medicine dropper and

tightly insert the stopper and long jet-tube. The water will slowly rise until it reaches the top of the tube, when it will come with a rush filling the flask with a beautiful fountain and at the same time the color of the water will change from red to blue. Ammonia is so soluble in water that as soon as the fountain starts a vacuum is created and the water continues to rise until the flask is nearly or quite full. Since the ammonia water is a base it changes the red litmus to blue.

These by no means exhaust the number of magic experiments that will appear in this book. Others will follow in the regular work of experimentation and special tests.

CHAPTER 5

PRIESTLEY AND SCHEELE

PRIOESTLEY and Scheele. When we hear these two names, we at once think of oxygen, for they were the discoverers of this important gas. And they were home-laboratory workers. Most of the early scientists were. High school and college laboratories did not then exist. And science in those days presented a veritable paradise of undiscovered possibilities. Most of the elements and their compounds were unknown. Almost any persistent experimenter might blunder onto some new discovery. But Priestley and Scheele were not blunderers. They were two of the most skillful chemists of their time. Priestley was an Englishman and Scheele a native of Sweden.

Priestley was forty before he began to experiment much. He was a clergyman who became interested in the study of gases. Many of the present pieces of gas apparatus were devised by him. One day he placed some red powder, now known as mercuric oxide, into a glass container and began to heat it. To obtain heat in those days was a difficult matter. Gas and the Bunsen burner were unknown, and the spirit lamps crude and inefficient. Priestley employed a large twenty-inch burning glass and focused the rays of the

sun upon his apparatus. The composition of this powder was unknown to him. Presently he noted that the contents of the tube turned dark in color. Then a silvery deposit appeared on the cooler portion of the tube and a gas began to bubble up through the water of his pneumatic trough. Surely he was making a discovery. He collected some of the gas. Finally he obtained a large bell-jar of it. To his surprise and delight, he found that substances burned in it with wonderful brilliancy. A mouse placed in the gas became very lively, and he himself was greatly invigorated by breathing it. Here was a new element, and Priestley immediately published his discovery. This was in 1774.

But at least two years earlier Scheele had also discovered this same gas but did not publish the fact until after Priestley had made his announcement. Therefore both may lay equal claim to this honor. Scheele was a poor apothecary's apprentice who worked throughout his short life of forty-three years in what we should call now a crude home laboratory. But he made remarkable discoveries. He discovered the poisonous gas chlorine and its bleaching properties. He perfected new methods of analysis and became the first great master of qualitative analysis. He carried out brilliant researches in organic chemistry and may be regarded as the founder of this important branch of the subject. Before he died he had been honored by the leading scientific societies of Europe and every school boy is now familiar with his name.

It is interesting to know that Priestley passed the last ten years of his life in this country and lies buried on the banks of the Susquehanna. Some of his original apparatus is now in the museum of Dickinson College at Carlyle, Pennsylvania.

CHAPTER 6

OXYGEN AND HYDROGEN

OXYGEN and hydrogen are two of the most important and widely distributed elements with which you will have to deal. The food you eat, the water you drink, and your own bodies contain both of these elements in combined form. The air is one-fifth oxygen, and all living things are largely compounds of oxygen and hydrogen together with carbon.

And right here we must distinguish between an element and a compound. *An element is a substance that cannot be changed by any known means into anything different from itself.* Of course radium does change into other elements, but we can do nothing to influence the process. There are something over 80 known elements, and each one has properties, or characteristics, which make it different from every other element. For instance hydrogen burns and oxygen supports combustion. Gold is yellow and silver is white. And no matter where an element is found its properties are always the same. But *a compound is a chemical union of two or more elements in which the elements have lost their characteristic properties and can be separated only by chemical means.* Water is a chemical compound of hydrogen and oxygen. You never would

suspect that water may be decomposed into two gaseous elements, one of which will burn and the other support combustion. The hydrogen and oxygen in uniting to form water have lost their own characteristic properties and have formed a totally new substance. To separate this new substance into its elements some kind of chemical means must be employed. We shall see what is meant by chemical means as we proceed.

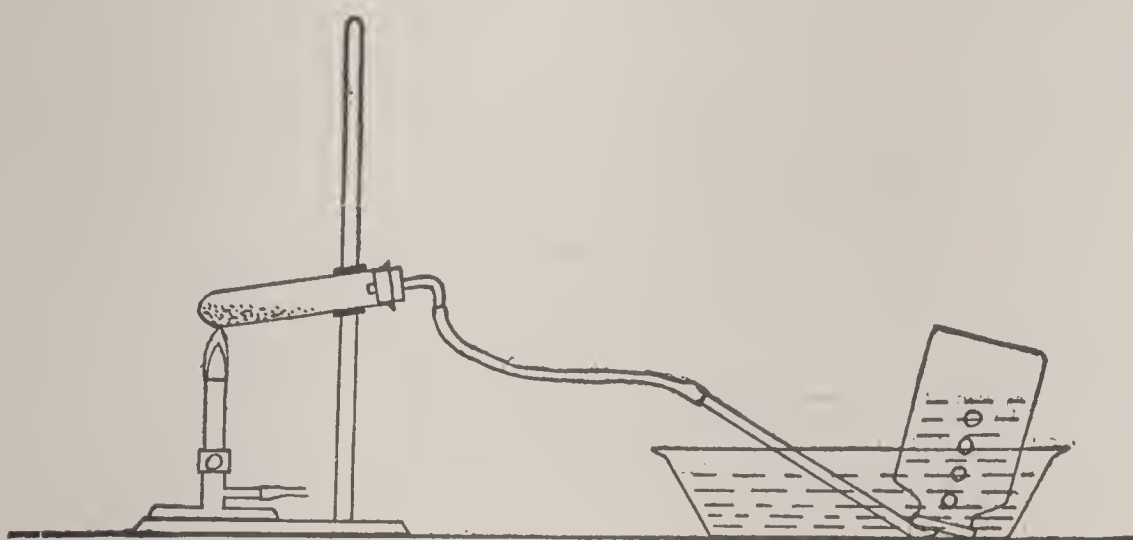


FIGURE 5.

Apparatus for the preparation and collection of oxygen.

Preparation of Oxygen.—Oxygen is the most abundant element known. Forty-seven per cent of the earth's crust is oxygen and eight-ninths of the ocean. It exists in many compounds, and some of them give it up readily. We shall select for our purpose in preparing it potassium chlorate.

Mix thoroughly two teaspoonfuls of powdered potassium chlorate with about one fourth that bulk of manganese dioxide. If you have a balance and weights, you may make a mixture of 8 grams of the chlorate and

6 grams of the dioxide. Place this mixture in a test tube fitted with a one-hole rubber stopper and delivery tube. For a pneumatic trough you may use a large basin (Fig. 5). Fill it nearly half full of water and invert in it a half dozen bottles full of water. To do this cover the bottle with a glass plate, quickly invert it, and when the mouth is beneath the water in the basin remove the plate. When all is ready, light your Bunsen burner or alcohol lamp and begin gently to heat the test tube, moving the flame about so as not to heat the mixture all in one spot. It will be best to have an assistant in this experiment. Let him tip one of the bottles slightly with one hand and with the other slip the end of the delivery tube beneath its mouth. Presently the gas will begin to come very rapidly, and one bottle after another may be filled until they are all full. If the gas is then still coming, hold in it a glowing stick and note the brilliancy with which it burns. This is the test for oxygen. You will observe that the oxygen seems to be white, but this is due to the presence of vapor and it will disappear upon standing over water for a short time.

Now why did we add the manganese dioxide? To determine this heat a little potassium chlorate by itself in a test tube, holding it with a test tube holder over the flame. The substance will melt and then boil. At that point thrust into the tube a glowing splint. It will burst into flame showing the presence of oxygen. Now allow the tube and contents to cool until oxygen just fails to be given off. Then add a pinch of manganese

dioxide, and immediately oxygen in large volume will escape. Don't you see, manganese dioxide lowers the temperature and increases the rapidity with which oxygen will be liberated?

Oxygen from Sodium Peroxide.—Oxygen may also be very readily prepared by the action of water on sodium peroxide. Place some of this powder which comes in a sealed can and must be kept sealed, in your gas generator. (See figures of apparatus in chapter 1.) Connect to it your delivery tube, the same as that used in the previous experiment and consisting of about two feet of red rubber tubing attached to a piece of glass tubing having its end slightly curled up. Through the thistle tube pour some water, being sure that the end of the tube nearly touches the bottom of the generator. Immediately the oxygen will begin to come off at a very rapid rate and may be collected over water in the usual way.

There are other methods of preparing oxygen, but these two are the simplest.

Burning in Oxygen.—You do not need to be told that all burning, or combustion, is due to the union of oxygen with other substances. To discover for ourselves the brilliancy of combustion in pure oxygen, we shall perform the following experiments.

Burning Steel Wool.—Let us start off with a real spectacular display. Turn up the end of a stout wire hook-shape and fasten to it a bunch of steel wool. If you do not have steel wool, unravel some picture cord wire and make it. Slip a glass plate beneath the mouth

of one of your bottles of oxygen and invert it upon the table, leaving it covered with the plate. Have at hand a small heap of powdered sulfur. Heat the wire in the Bunsen flame and touch it to the sulfur. Some of the sulfur will take fire and cling to the wire. Immediately remove the glass plate from the jar and thrust into it the ignited wool. The combustion will be of dazzling brilliancy accompanied by a shower of sparks. If this can be done in the dark, the vividness will be much increased.

Combustion of Magnesium.—Wind 4 or 5 inches of magnesium ribbon about a length of stout iron wire, allowing the lower end of the ribbon to project a little. Ignite this end by touching it to the Bunsen burner flame and quickly thrust it into a bottle of oxygen. Again there will be a flash of blinding light.

Combustion of Phosphorus.—Line the bowl of your deflagrating spoon with asbestos paper and place in it a little *red* phosphorus. Ignite the phosphorus in the flame and lower the spoon into a bottle of oxygen. The combustion is brilliant and a white cloud of phosphorus pentoxide will fill the jar.

Burning of Powdered Charcoal.—In the bottom of a dry pint milk bottle place a layer one-fourth inch deep of finely powdered willow charcoal which has been previously dried in an oven. Prepare a fresh mixture of potassium chlorate and manganese dioxide and generate more oxygen, but this time thrust the delivery tube to the bottom of the bottle and direct the stream of gas so that it will produce a cloud of charcoal dust.



Burning steel wool in pure oxygen.

Then remove the delivery tube and immediately apply a flame by means of a long wax taper. The combustion will be explosive, but there will be no danger from it.

Burning of Powdered Iron or Zinc Dust.—Repeat the previous experiment, but substitute powdered iron or zinc dust for the charcoal. In order to protect your eyes from the blinding light, it will be necessary to wear colored glasses.

Combustion of Sulfur.—Again line the bowl of your deflagrating spoon with asbestos and place in it a piece of sulfur about the size of a pea. Ignite it and lower the spoon into a bottle of oxygen. Note the bright blue flame and how much more brilliant the combustion is in pure oxygen than it is in air. Cautiously smell of the gas formed in the bottle. It is sulfur dioxide so much used for fumigation.

The above experiments are examples of rapid combustion, but slow oxidation unaccompanied by light is taking place about us all the while. The heat of your body is being maintained by oxidation. All forms of decay, the drying of paint, the rusting of iron, the corrosion of metals, the bleaching of colored fabrics, the changing of cider to vinegar, and the antiseptic action of hydrogen peroxide are all examples of oxidation.

Spontaneous Combustion.—Fires frequently start without any apparent cause. Wet hay in a mow, piles of coal dust, heaps of oily waste, rags saturated with paint or turpentine, and substances with a great affinity

for oxygen take fire spontaneously. A painter once told me that an oily rag, which he had stuffed into his pocket, took fire and burned his leg. I once had a quite serious fire in my laboratory from the spontaneous combustion of sticks of yellow phosphorus, which had become exposed to the air.

To illustrate this action place a few teaspoonfuls of carbon disulfide in a small glass-stoppered bottle and drop into it a piece of yellow phosphorus about the size of a pea. The phosphorus will quickly dissolve in the liquid.

Two Cautions: Keep carbon disulfide away from a flame. Always cut yellow phosphorus in a basin under water, always keep it covered with water, and never touch it with the fingers except under water.

If you should ever get a phosphorus burn immediately apply a solution of potassium permanganate and continue to apply it at frequent intervals until the burn is healed.

Now to come back to our illustration of spontaneous combustion, pour a little of the solution of phosphorus in carbon disulfide upon a piece of absorbent paper such as a filter paper and, holding it with pincers or tongs, wave it back and forth through the air. In a few moments the paper will take fire. What has happened? Simply this—the carbon disulfide has evaporated leaving a finely divided layer of phosphorus on the paper. The phosphorus immediately begins to unite with the oxygen of the air and as it does so generates enough heat to bring it to its kindling tem-

perature. When this point is reached nothing can prevent it from bursting into flame.

In this experiment be very careful not to get any of the solution on your hands or clothes.

How Gasoline Explodes.—Find a molasses can with a tight fitting cover. In the side and near the top punch a hole about a half inch in diameter. Place the can on a ring-stand. Set under it a candle and at the side a tall candle so that the top of it will come just under the hole that you have made. Now put six drops of gasoline or ether in the can and fit the cover on tightly. First light the candle at the side, then the one under the can, and step back. Soon an explosion will occur.

As the gasoline vaporizes and mixes with the air an explosive mixture fills the can, which is ignited by the candle at the side.

How Illuminating Gas Explodes.—In the top of another quart-size molasses can punch a hole about three-eighths of an inch in diameter. Make a second hole of the same size and near the bottom of the can. Through the hole at the bottom thrust a rubber tube connected to a gas jet. Allow the gas to run for about 20 to 30 seconds, filling the can. Then turn off the gas, quickly remove the tube, and light the gas at the hole in the top of the can.

The flame will grow smaller and smaller, but do not think it has gone out. Keep away and presently an explosion will occur. As the light gas escapes at the top and burns, air enters through the hole in the side

and, when the can has become full of an explosive mixture, the flame strikes down and ignites it. This is the cause of the striking back of a gas burner and the frequent explosions that occur wherever gas escapes.

A Dust Explosion.—Obtain another can similar to the other two but somewhat larger if possible. Punch a round hole in the middle of the bottom of it and insert through this from the inside your funnel, pushing the stem down so that it projects underneath. Attach to the stem a 2 or 3 foot length of rubber tubing. Rest the can upon a ring-stand support. Place inside the can beside the funnel a candle about an inch long and light it. Have at hand an ounce of lycopodium powder, which you will be able to obtain from a druggist, or thoroughly dried powdered charcoal. Raising the free end of the rubber tubing, so that the powder cannot run out, pour a teaspoonful of it into the funnel and press the cover tightly into place. Then taking the tube in your mouth blow into the can with a strong quick breath, being careful to keep back as far as possible.

The explosion is due to the sudden ignition of the mixture of dry combustible dust and air. A chance spark or flame in a dusty grain elevator or in a plant for grinding materials such as paint pigments frequently cause explosions of great violence.

Oxidation is one of the fundamental types of chemical change. We shall run across it time and again in our work, and before long it will become as familiar as an old friend.

HYDROGEN

Hydrogen is quite as interesting an element as oxygen. It was discovered by Sir Henry Cavendish in 1766, and its name, meaning "water former," was given to it by Lavoisier. As a constituent of almost all living organisms, as well as water, it is of tremendous importance. It occurs also in petroleum, natural gas, and coal tar. Until displaced somewhat by helium, it was the chief gas used in filling balloons and airships.

Preparation of Hydrogen.—Arrange your gas generator and pneumatic trough exactly as you did in the experiment with nitric acid and copper rivets in the chapter on Chemical Magic. See Fig. 4. But this time place in the generator granulated zinc or sheet zinc cut into small squares.

Upon this zinc you will pour dilute sulfuric acid, and I must tell you how to prepare this solution. First, always remember that sulfuric acid must be *poured into water* and never water into the acid. In as large a beaker as you have place 300 or 400 c. c. of water and pour into it slowly and with constant stirring with a glass rod one-sixth that volume of the strongest sulfuric acid. The solution will be quite hot and before you use it let it cool. Then pour it into a labeled glass-stoppered bottle.

The gas will be collected over water, as oxygen was. Have everything in readiness and be sure that the joints in your generating apparatus are all perfectly

tight. Then pour a little of the acid onto the zinc, enough just to cover it well, and take care to have the thistle tube dip beneath the surface of the acid. Keep your Bunsen burner at a safe distance from the apparatus. As soon as the acid comes in contact with the zinc, hydrogen gas will begin to be liberated, but do not collect it in the bottles just yet. The generator is at first full of air, and hydrogen and air form an explosive mixture. Over the delivery tube hold an inverted test tube, and, since hydrogen is very light, it will rise and displace the air. Keeping the test tube mouth down bring it to a flame. A sharp report will be heard. This is because the air has not all been swept out of the generator. Repeat the operation, but do not bring the test tube immediately back to the delivery tube, for a flame plays about its mouth for a few moments, and, should you do so, the flame would strike back into your generator and the mixture would explode, driving the thistle tube to the ceiling.

When the gas burns quietly in the test tube on presenting it to the flame, you may begin to collect the gas in the bottles. As the action slows down add more acid. Collect several bottles of the gas and leave them inverted over water until you are ready to use them.

Combustibility of Hydrogen.—Holding an inverted bottle of hydrogen in the left hand, thrust up into it a lighted taper. Note that the hydrogen ignites at the mouth burning with an almost colorless flame. Also observe that the taper is extinguished as it is thrust

into the gas, but relights as it is withdrawn from the bottle. This extinguishing and relighting of the taper may be repeated over and over, showing that hydrogen is combustible but does not support combustion.

Hydrogen Lighter than Air.—Holding a bottle full of air in the left hand and a bottle of hydrogen in the right, slowly bring the bottle of hydrogen mouth upward beneath the bottle of air. Then present each bottle in succession to the flame. A sharp report will result in each case, showing that hydrogen must have risen upward into the second bottle. Were hydrogen a colored gas, this movement could be very readily seen.

Affinity of Hydrogen for Oxygen.—In a dry glass test tube place some copper oxide, preferably of the wire form, and clamp the test tube to your ring-stand support so that the open end inclines slightly downward. Get your hydrogen generator into action again and be sure the gas is coming free from air. Then pass the delivery tube nearly to the bottom of the test tube (Fig. 6). Let the gas run for a few moments to remove the air from the tube, and then heat the tube directly under the copper oxide, moving the flame back and forth as you do so. In a few moments steam will begin to issue from the tube and water will condense on the cooler portions. At the same time bright copper will form in the test tube. If you allow the hydrogen to continue to pass into the tube as the copper cools, it will remain bright.

The alchemists would have said that this action

represents the transmutation of a black brittle substance into copper. We now know that it is due to the union of hydrogen with the oxygen of the copper oxide. This kind of a chemical action we call *reduction* and hydrogen is a *reducing agent*. Oxidation and reduction are twin processes. They go hand in hand. We may have oxidation without reduction, but not the reverse.

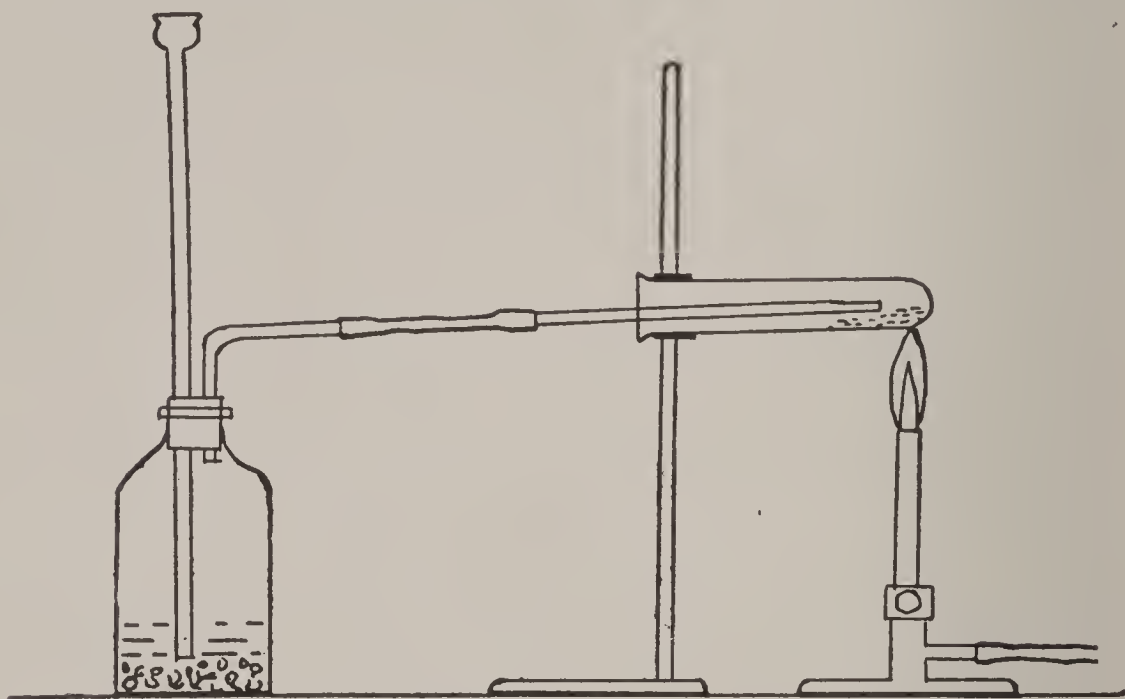


FIGURE 6.

Reduction of hot copper oxide to metallic copper with hydrogen.

A Safe Explosion.—A perfectly safe explosion of hydrogen and oxygen and one that will afford you plenty of amusement may be easily carried out. Obtain from a druggist a bottle holding from one to two quarts and having a ground glass stopper. It will be necessary to cut the bottom off this bottle, but the following process will do it. Soak a hemp string in kerosene and draw it tightly about the bottle close to

the bottom. Light the string and when it is through burning thrust the bottle into cold water, the colder the better. The bottom will usually crack off along the line of the string.

Now using a deep jar for a pneumatic trough, fill the bottle with water and displace it with hydrogen from your generator. Have ready a one-hole rubber stopper fitted with a short piece of glass tubing, as large in diameter as the stopper will permit of. When the bottle is full of hydrogen, rest it on two small blocks of wood so as to raise it about an inch from the table, quickly remove the glass stopper and insert in its place the rubber one. As you do so, direct your assistant to light the hydrogen escaping from the tube in the stopper. Then step back.

The flame will grow small and may seem to disappear but do not become impatient. Presently it will strike down into the bottle and the explosive mixture will ignite with a very sharp report. As the hydrogen escapes, air enters from the bottom, forming the combustible mixture.

Hydrogen from Water.—If your supply of chemicals contains metallic sodium, you should know that this substance must be kept under kerosene, for reasons that you will presently understand. It usually comes from the supply house in a sealed container. Open it and place the metal sodium in a bottle of kerosene.

For this experiment cut off a piece about the size of a pea, pare off any yellow crust that may be on the outside, and wipe off the kerosene with filter paper.

Wrap the sodium in a piece of sheet lead, leaving the tip exposed. Invert a test tube filled with water in your pneumatic trough and, grasping the sodium capsule with pincers, quickly thrust it beneath the mouth of the inverted tube. Immediately the sodium will begin to decompose the water and the test tube will fill with hydrogen. Present the tube to a flame and the gas will burn with a slight explosion.

Hydrogen from Steam.—Some metals like sodium

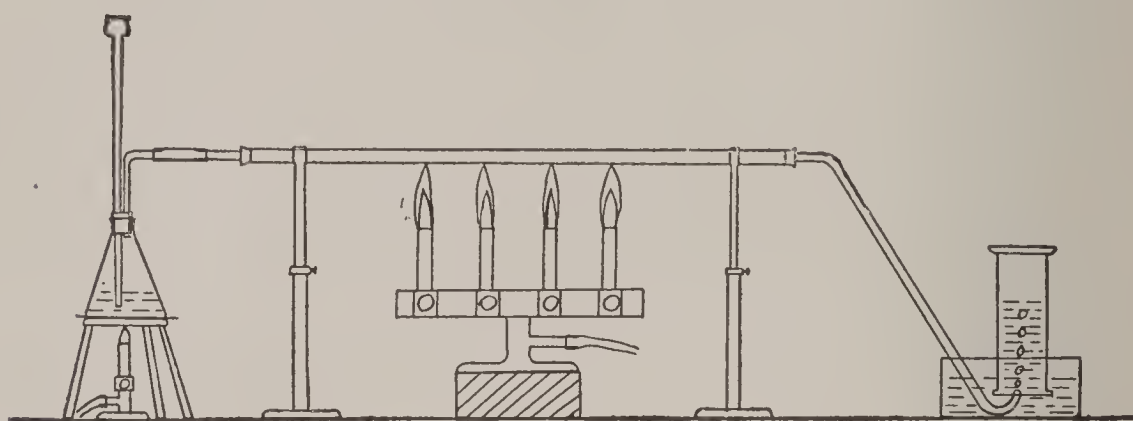


FIGURE 7.

Preparation of hydrogen by passing steam over hot iron.

will liberate hydrogen from water at ordinary temperatures. Others will do so only when heated.

For this experiment obtain a 24-inch length of iron pipe. Fill it with small iron nails and mount the tube on two ring-stand supports. Place in each end either a one-hole cork stopper, or better a red rubber stopper, each stopper having inserted in it a short length of glass tubing. Connect to one end by means of a short piece of rubber tubing a flask containing a small quantity of water and mounted on a ring-stand over a Bunsen burner. To the other end attach your delivery

tube and put in position the pneumatic trough and inverted bottles of water. In order to heat the iron pipe you will need what is called a 4-tube burner, that is, 4 Bunsen burners mounted so as to be fed from a single gas connection. You will find this piece of apparatus useful in other parts of the work too. (See Fig. 7.)

Now light the 4-tube burner and bring the iron pipe to a red heat. Then bring the water in the flask to a boil and pass steam over the red-hot iron. A very rapid stream of hydrogen will be liberated from the steam, which may be collected in the usual way.

Many of you will not have the facilities for this experiment, but it will be interesting for those who have.

Blowing Soap Bubbles with Hydrogen.—Between the delivery tube and the hydrogen generator place a short length of glass tubing containing a plug of cotton wool. To the end of the delivery tube attach a thistle tube or clay pipe. Prepare a strong solution of soap, as you have doubtless done many times before, but put into it this time a considerable quantity of glycerin.

Now put the generator into action by pouring dilute sulfuric acid onto the zinc. Before you blow any bubbles test the gas to be sure that it burns quietly and is free from air. Dip the thistle tube into the soap solution and then almost immediately direct it upward. As the bubble increases in size give the tube a slight twist and it will release itself, rising rapidly to the

ceiling. It will also be interesting to light these bubbles as they are released.

Similar bubbles may also be blown with ordinary illuminating gas.

What Forms When Hydrogen Burns?—Probably you have discovered the answer to this question already, but to confirm it thrust the delivery tube of your generator into a dry test tube held mouth downward. After the hydrogen has had time to drive out the air, hold the tube to the flame and note that as the hydrogen burns a film of moisture covers the sides of the tube. Water is the product.

CHAPTER 7

SIR HENRY CAVENDISH

SIR HENRY CAVENDISH was one of the queerest scientists who ever lived. He was one of the wealthiest men in England, but he cared nothing for his money. His sole pleasure in life came from experimenting. And probably no man of his time was more expert as a glass-blower and an analyst than he. He did remarkably accurate work for those days, but it was entirely for his own satisfaction. He did not want anyone else to know about it. Now, we should call him selfish and eccentric.

But we remember Cavendish chiefly for his discovery of the element hydrogen. Hydrogen, is the lightest known substance. It is the gas with which we fill balloons and airships. And shortly after Cavendish's discovery it was used to lift into the air some of the first balloons. Cavendish found that this gas burns with a very hot flame and in doing so produces water. Because it all seemed to disappear on burning, he called it pure phlogiston. And although he lived at the time of Priestley, Lavoisier, and Scheele, Cavendish never accepted the new explanation of combustion. He clung to a belief in some sort of fire-stuff that disappeared on burning.

After a time, in spite of all Cavendish's opposition, the knowledge of his discoveries became known to other scientists. Then they sought to honor him. But Cavendish would have none of it. On one occasion he fled precipitately from the back door of the banquet hall and left those who had gathered to do him honor to go on without his presence. Even in death he savagely dismissed his servant that he might die alone. But although this old chap was exceedingly unlikable in his disposition, we must honor him today because of his great discoveries.

CHAPTER 8

THE ATMOSPHERE

WE live at the bottom of a great ocean. It is as truly an ocean as are the immense bodies of water that bathe the continents. And we are as uneasy out of it as a fish out of water. In fact we cannot live without it. The air forms a very interesting subject for study. Possibly we may learn some things about it that you do not already know. Let us see.

Does the Air Have Weight?—Place in one side of your hornpan balance a burned-out electric lamp. Just exactly counterpoise it by placing lead shot and bits of paper in the other pan. Then light your Bunsen burner and adjust the flame so that it is about an inch and a half high. (Of course an alcohol lamp may always be substituted for a Bunsen burner.) Holding the lamp close to one side of the flame, with your blow-pipe direct the flame against the glass. In a few moments the glass softens and a hole is forced through it. You will also hear a sound like the escape of air through a small opening.

Now place the lamp back in the balance pan. Note the increase in weight. An electric lamp bulb contains a vacuum, and, when you make a hole and admit air, the vacuum is destroyed and the bulb gains in weight.

You may also prove this fact that air has weight with a football bladder. Attach it to one side of the balance and counterpoise it with lead shot on the other side. Then blow the bladder up and attach it to the balance again, using the same amount of string to fasten it that you did the first time. Since the quantity of air weighed in this experiment is much greater than in the other, the weight change will be much more pronounced.

Presence of Oxygen in Air.—Of course the fact that substances burn in air would seem to indicate that air contains oxygen, but we need better proof. Obtain a quart fruit jar with a tight-fitting top of the clamp-fastening type. *Observing the precautions about the use of yellow phosphorus*, cut a piece under water about the size of a pea, wipe it very gently with a filter paper, being careful not to touch it with the fingers, and place it in your porcelain crucible set upon the upturned cover of the fruit jar. Have the jar fitted in advance with two new rubber rings well greased on both sides. Invert the jar over the cover and clamp it securely into place. Keeping the jar still inverted immerse the end containing the phosphorus in a basin of hot water for a few minutes. Presently the phosphorus reaches its kindling temperature and begins to burn, forming a dense white cloud inside the jar.

After the action has stopped and the jar has cooled for *fifteen minutes*, thrust the end containing the crucible into a basin of cold water. Unfasten and pry off the top being careful to keep that end of the jar all

the while immersed. As the top is removed, water will rise in the jar. Why?

Slip a glass plate under the mouth of the jar and quickly invert it upon the table, keeping the water in the jar. Shake the water about gently so as to dissolve the white substance formed in the combustion. Thrust a burning splint into the air remaining in the jar. Note that the air no longer supports combustion. It must have lost something.

Now pour into the water in the jar blue litmus solution. The red color appearing shows the presence of an acid. Neither the air nor the water contained an acid at the start. It must have come from the change. Chemists know that this new compound is phosphoric acid and that it contains oxygen. Therefore the oxygen must have come from the air.

There is *another and simpler way* of demonstrating the presence of oxygen in the air and one that you will understand more easily. Fill your porcelain crucible with short lengths of about No. 22 copper wire. Using the ring-stand support and a clay triangle (see Forms of Apparatus) place the crucible over the flame and gradually heat it to as high a temperature as possible. Occasionally stir the wire so as to bring all of the pieces in contact with the air. After about 30 minutes turn off the flame and allow the crucible to cool. The copper is no longer bright and flexible but black and brittle. It has undergone some chemical change. Let us discover what it is.

Place some of the new copper compound in a test

tube clamped to a ring-stand, just as you arranged the apparatus in the reduction experiment with hydrogen. Set up your hydrogen generator and with zinc and sulfuric acid get it into action. Always remember to use dilute sulfuric acid, and I will tell you the secret of starting the action when the gas is slow in coming. Add through the thistle tube a few c. c. of copper sulfate solution. It will work like magic.

When you have tested the hydrogen and found that it is free from air, that is, a test tube of it burns quietly without explosion, heat the test tube containing the copper compound and thrust the delivery tube up into it. In a few moments steam and water will come from the tube and bright copper will be left.

Now must we not conclude that the new copper substance contained oxygen? And from where did it get the oxygen, if not from the air? I think this proof will afford you considerable satisfaction. The explanation looks simple now, but a century and a half ago it would have been a profound mystery.

Another New Substance.—In this experiment with hydrogen and in the preceding ones, you must have noticed that the zinc disappears and the acid is used up. Your generator is not a perpetual motion machine. It will not go forever. To prove that a new substance forms in the generator, filter some of its contents into your evaporating dish or onto a watch glass. Evaporate this solution over steam. This can be done by resting the dish or watch glass on the top of a “lipped” beaker containing water and mounted

over your burner (Fig. 8). As the water boils, the escaping steam will evaporate the solution to dryness. In this way you will find that a new compound forms in the generator. It is a white substance called *zinc sulfate*.

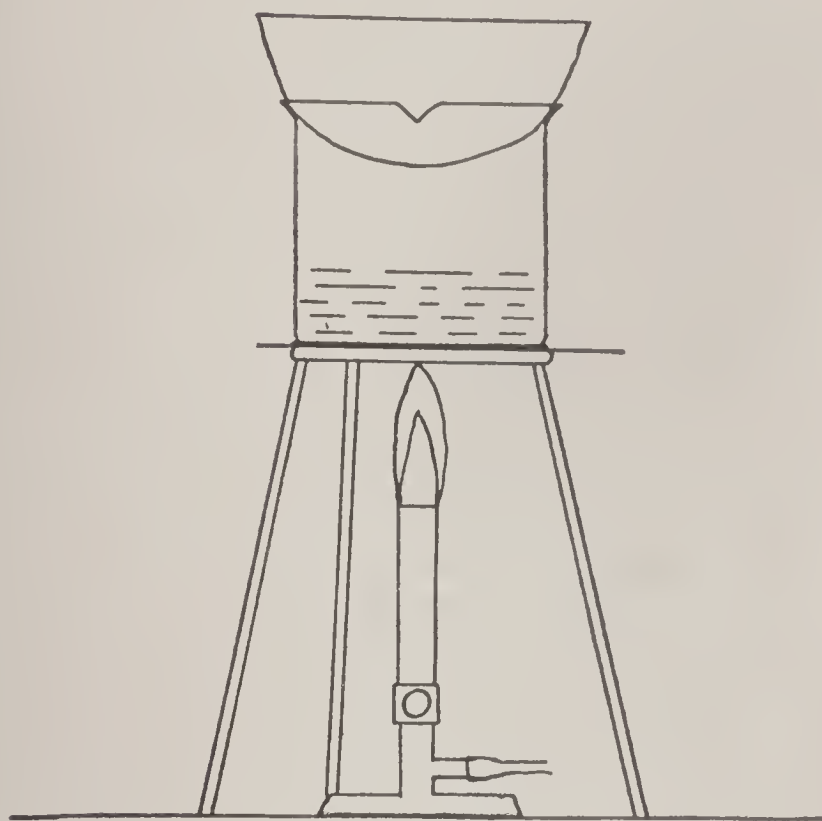


FIGURE 8.

A water bath for evaporating by steam.

Carbon Dioxide.—You have often heard of carbon dioxide. The gas that frequently escapes from your stomach after drinking soda water is carbon dioxide. It is generated in processes of fermentation and decay. It is produced by yeast and baking powders in bread making and causes the bread to rise. It issues from the earth and is a product of important chemical changes.

Is carbon dioxide present in the air?—Set a small beaker of limewater aside for a few hours. A white crust will form on the surface. With a bicycle pump force air through a tall cylinder or deep bottle of limewater for 20 minutes. Note the milky appearance of the limewater. As we shall see, this turning of limewater milky is a test for carbon dioxide.

Preparation of Carbon Dioxide.—Place in your hydrogen generator some lumps of marble and arrange it just as you would for the preparation and collection of hydrogen. Being sure of course that the thistle tube reaches to the bottom of the bottle, pour through it upon the marble just enough water to cover the pieces and then add strong hydrochloric acid. Immediately a vigorous action will be set up and a rapid stream of gas will escape through the delivery tube, which may be collected over water in the usual way. Collect a number of bottles of the gas. It is carbon dioxide.

Slipping a glass plate beneath the mouth of one bottle invert it upon the table. Pour into it a little limewater and shake. You will observe the same milky appearance obtained from pumping air through limewater, except that it is much deeper.

Pour another bottle of the gas downward upon a lighted candle. By this experiment you prove that the gas is heavy and that it does not support combustion. Into another bottle thrust a burning splint. Again you obtain more proof of the character of this gas.

Again start the action in the generator and pass the



Generating carbon dioxide and collecting it over water.

gas through blue litmus by dipping the delivery tube into some of the solution placed in a test tube. What does the color change show you about this gas? It is sometimes called carbonic acid gas. With water it gives the acid. Soda water is nothing but this gas dissolved in water under pressure.

What are the causes of carbon dioxide in the air?—Through a glass tube having about an inch of the upper end bent slightly, bubble your breath through a little limewater in the bottom of a test tube. The tell-tale color denotes the presence of the gas. Respiration is then one source of carbon dioxide in the air.

Lower a burning splint into a bottle of air. After a few moments withdraw it and pour limewater into the bottle. Shake the contents and note the evidence of carbon dioxide. Combustion is another source.

By means of a wire attached to a lighted candle lower it into a quart bottle and cover the bottle tightly with a glass plate. Note how long the candle continues to burn.

Then fill the same bottle with water and invert it in your pneumatic trough or in a basin of water. Through a rubber tube dipping beneath the slightly upturned mouth of the bottle, blow into the bottle and with repeated breaths fill it with air from your lungs. Remove the bottle and lower into it a lighted candle, covering it as before with a glass plate. Again note the length of time the candle burns. What difference is there between fresh air and air expelled from the lungs?

Carbon dioxide is not a poison. Its presence in the air is sometimes an indication of the state of purity of the air, but in itself it is not harmful. On the other hand it has a number of important uses. Plants breathe it. It is a valuable fire extinguisher. When great clouds of it are produced by smudge pots and made to hover over the orchards of the West it prevents frosts and saves millions of dollars' worth of fruit. Without it you would be deprived of the pleasures of soda water.

Moisture in the Air.—In a bright metal cup place a little ether and blow into it through a glass tube. In a few moments a film of moisture will appear on the outside of the cup. From where did it come? Evaporation of course produces cold, as we say, but it more properly absorbs heat. If it did not, evaporation would not take place. This heat comes from the air and quickly cools it to its dew point, the temperature at which condensation of moisture begins.

If you have no ether, stir into some water placed in the cup a little cracked ice. This will also precipitate moisture on the outside of the cup.

Your breath is constantly adding moisture to the air. Confirm it by breathing onto a cold window pane. Another source of moisture is combustion. Place your wash bottle full of cold water but dry on the outside on the ring-stand and light under it the Bunsen burner. Immediately a film of moisture will spread over the flask and often drops of water will run down its sides. From where does it come? At first this always

seems a perplexing problem. But when we remember that a large percentage of illuminating gas or alcohol, if you use an alcohol lamp, is hydrogen, which produces water upon burning, the answer is simple.

Nitrogen in the Air.—Although nearly four-fifths of the air is nitrogen, this gas by itself is uninteresting. There is nothing we can do with it. It does not enter readily into combination with other elements. After we had removed the oxygen from the air with phosphorus, you will recall that we thrust a burning splint into the remaining gas, but nothing happened, except that the splint went out. This gas was nitrogen and we learned that it neither burns nor supports combustion. But in its compounds nitrogen is a most fascinating element. Because of its weak chemical attraction for other elements, it plays the leading rôle in the destructive power of high explosives. And its compounds have many other uses. One of the three chief kinds of foods must contain nitrogen. We shall run across it again a number of times before we are through with this work.

Of course the air also contains about one per cent of the so-called rare elements, discovered by Sir William Ramsay, and like nitrogen exceedingly inactive, so much so that they form no compounds at all. The most interesting one of them is *helium*, which is being used now as a substitute for hydrogen in filling airships. It is not quite as light as hydrogen, but it will not burn, and that avoids the frightful accidents

that sometimes occur when a spark from the motor pierces the gas bag.

I am sure you will agree now that the atmosphere possesses much of genuine chemical interest.

CHAPTER 9

ANTOINE LAURENT LAVOISIER

IN Lavoisier, a brilliant French chemist, we have another who at an early age established a well equipped private laboratory and turned his attention to science. That was the day of the home-laboratory worker. Lavoisier was born in 1743, the son of wealthy parents. He was able to put into his laboratory everything that science and money could provide. But best of all he had a keen mind and a wonderful ability to interpret, or explain, the experimental facts which he observed.

Scientists in those days had a very queer notion of just what happens when a substance burns. And if we were not the heirs of a century and a half of chemical investigation, we would still have the same strange idea. Maybe you do not know the cause of burning even yet. We call it *combustion* now. Up to the time of Lavoisier chemists supposed that in combustion an imaginary substance called "*phlogiston*" is given off. What more natural, for something surely disappears? Coal for instance seems to be almost entirely consumed. But Lavoisier attacked this problem with a balance. He weighed everything entering into the chemical change and all the products. Nothing

escaped him. And he found that when a substance is heated or burned in the air the products of the change weigh more than the original substance itself. Even though a candle seems to disappear, the products of its combustion weigh more than the candle did itself. Clearly something from the air must unite with the candle. But what this something might be puzzled Lavoisier. He performed many experiments with the balance and always obtained similar results.

Then one day Lavoisier learned of Priestley's discovery of oxygen; although it was not called oxygen then, but "dephlogisticated air." All gases were called "airs" in those days. And this bit of scientific news from across the Channel gave Lavoisier just the clew he needed for the explanation of combustion. He saw at once that oxygen is the element in the air that enters into combination when a substance burns. He also showed that the rusting of iron and many other chemical changes are due to the union of substances with this newly found element in the air, for Priestley and Scheele had shown that it exists in the atmosphere as well as in many chemical compounds.

Because Lavoisier believed oxygen a necessary constituent of all acids, he gave it its name, which means "acid-former." We now know this idea to be incorrect, but the name will always remain. Lavoisier lost his life in the French Revolution and science thereby one of its most distinguished men.

CHAPTER 10

ACIDS, BASES, AND SALTS

ACIDS, *bases*, and *salts* are the three fundamental types of compounds with which chemistry has to deal. Without knowing it you are already familiar with many of these substances. One of the chief characteristics of an acid is its sour taste. You know the sour taste of vinegar, lemons, other citrus fruits, and plain soda water. You have drunk volumes of carbonic acid in soda water, but fortunately it is a weak acid and it also decomposes in your stomach into two harmless products. Ammonia, lye, and limewater are bases. And ordinary table salt, borax, washing soda, cream of tartar, baking soda, and blue vitriol are common salts.

There is a very close relationship between these substances, as we shall see. In order to test their properties you will need two or three solutions of indicators. Litmus, methyl orange, and phenolphthalein are the three most common indicators. The first comes in the form of small cubes, and the other two as powders. To prepare litmus solution place a small handful of the cubes in a beaker, add 300 or 400 c.c. of water, and bring the mixture to a boil. Allow it to cool and

pour the clear solution into a stoppered bottle. Dissolve a small pinch of methyl orange in a little denatured alcohol and dilute with a considerable volume of water. Prepare phenolphthalein solution in the same way but do not add water.

Now acids and bases give characteristic colors with these indicators. You have observed the changing colors of the leaves in autumn; they are due to the formation of acids in the sap of the tree.

Color Effects of Acids and Bases.—Place a row of 8 small beakers or tumblers on your table. In every other beaker place an acid, using hydrochloric, nitric, sulfuric, and acetic. Only a very little of each will be required. In the remaining beakers place solutions of sodium and potassium hydroxides, ammonia water, and limewater.

(The sodium and potassium hydroxide solutions may be prepared by dissolving a 3-inch stick of each in 200 c.c. of water. Household ammonia will always be at hand, and limewater is prepared by shaking a handful of slaked lime with a quart of water and allowing the mixture to stand for a few hours. Then pour off the clear liquid.)

Now dilute the contents of each beaker with a little water from your *wash bottle*. (See Forms of Apparatus in first chapter.) Have at hand three small bottles containing the indicators. By means of a medicine dropper introduce one or two drops of litmus into each beaker. Note the characteristic red obtained with each acid and the blue with each base.

Litmus is the indicator commonly used to test for an acid or a base.

You will find it interesting to pour the contents of one beaker into that of another. In some instances the blue will predominate, in others the red. Why?

Rinse out the beakers thoroughly and arrange another series of acids and bases just as before. But this time use methyl orange indicator, introducing it with a fresh dropper. What colors do you obtain now?

Again repeat the experiment, using phenolphthalein indicator. You will observe that only the basic solutions give a color in this case.

Testing Common Things.—Squeeze out and filter into clean test tubes the juice of lemons, grape-fruit, orange, and a sour apple. In other test tubes dissolve with a little water small quantities of lye, baking soda, borax, and washing soda. Then in succession divide each solution into three portions and test them with the above indicators. Although the three last substances taken are salts, for reasons that you would not now understand, they give basic, or *alkaline*, reactions.

Shake some wood ashes in a test tube with water, let them stand for a while, filter and test for acid and base. Ashes are used to destroy the acid condition of old soil, or to sweeten the soil, as the farmer says.

Dissolve soap in water and test the solution with phenolphthalein. The cleansing action of soap is partly due to its alkaline reaction.

Preparing a Salt.—A salt is formed by the chemi-

cal union of an acid and a base, and the process is called *neutralization*.

Dissolve a piece of sodium hydroxide about as big as a pea in a test tube of water and pour half of the solution into an evaporating dish. Prepare a dilute solution of hydrochloric acid by mixing 10 c.c. of the strongest acid you have with 100 c.c. of water. Nearly fill a test tube with this acid and insert in the mouth of the test tube a cork having two small notches cut in opposite sides of it. This will afford you a good dropper.

The most convenient indicator for general use in your laboratory will be litmus in the form of red and blue litmus paper. This can be readily obtained, and you should cut it into small narrow strips.

On a small glass plate beside your evaporating dish containing the sodium hydroxide solution place a strip of each kind of paper. Prepare a glass stirring rod by fusing in the flame the ends of a piece of tubing. Then add the acid to the base with constant stirring, a few drops at a time and after each addition touch the end of the stirring rod to each kind of paper. At first the red paper will be turned blue, but after a time you will reach a point where the reverse change occurs. This means that you have over-stepped the *end-point*, or *neutral point*. The solution is now slightly acid. Therefore add a few drops of base from the tube in which you prepared it and bring the solution back to a faintly alkaline condition. You will then have reached a point where a single drop of acid will cause

the solution to turn blue litmus red, and a single drop of base will cause it to turn red litmus blue. The acid and base have now neutralized each other and you have in solution a salt. Dip the glass rod into the solution and touch it to your tongue. What is the salt?

Now evaporate the solution to dryness, using a small flame at the end so as to avoid spattering. You have prepared ordinary table salt, or sodium chloride, and if your chemicals were pure, the salt will be chemically pure.

There are many salts, and with any other acid and base in like manner you may prepare others.

Titration.—The process of neutralization is of great importance in analytical work. As carried out there, it is called *titration*. While we shall do no analytical work at this time, it will be interesting to learn the process.

For this work you will need two *burettes* holding 50 c.c. each. (See Forms of Apparatus.) In one, place a dilute solution of hydrochloric acid, and in the other, a dilute solution of sodium hydroxide. Mount these burettes on a ring-stand support by means of clamps (Fig. 9). Have at hand a small beaker, a stirring rod, and phenolphthalein indicator. On the table under each burette place a square of white paper. Place the beaker under the sodium hydroxide burette and run out about 10 c.c. of the base, allowing it to run down the stirring rod. Remove the drop clinging to the tip of the burette with your stirring rod and

wash down the rod and the drops spattered on the inside of the beaker with a stream of water from your wash bottle.

Now add a drop of phenolphthalein to the contents of the beaker and run in acid, at first rapidly and then

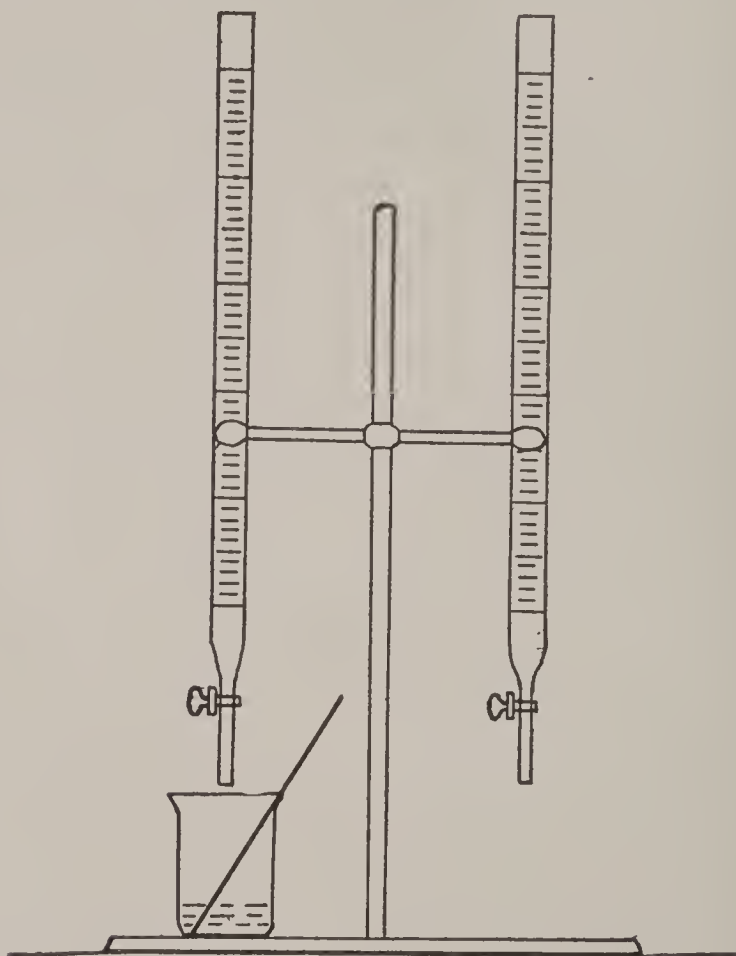


FIGURE 9.

Burettes set up for determining the strength of an acid or a base by titration.

drop by drop, until the pink color produced by the indicator is just destroyed. Again remove that last drop clinging to the burette and rinse down the rod and sides of the beaker. If the work has been done carefully, you have now reached a point where a single



Determining the strength of an acid by titration.

drop of either solution will produce the color change. This is the end-point. Taste of the neutralized solution. If you like, evaporate it, as before.

You are now getting to be a real chemist, for these are real chemical processes. An interesting additional experiment will be the titration of "white" vinegar with sodium hydroxide as the base. Ammonia and hydrochloric acid may be titrated, using either litmus or methyl orange as the indicator.

Preparation of Hydrochloric Acid.—As we have seen an acid combined with a base produces a salt. Therefore we ought to be able to obtain an acid from its salt, and that is exactly the method that we shall employ.

Mount a 500 c.c. flask on your ring-stand and fit it with a 2-hole stopper carrying a thistle tube and bent angle tube. Attach the delivery tube and pass it through a cardboard cover nearly to the bottom of the collecting bottle (Fig. 10). Place three or four tablespoonfuls of ordinary table salt in the flask and moisten it with a little water. Then add about 20 c.c. of the strongest sulfuric acid through the thistle tube and heat the flask gently with the flame. Be sure you have an asbestos gauze under the flask, and if possible place the apparatus in a good draft.

Now hydrochloric acid is a gas dissolved in water, and what you will obtain here is the gas. It will begin to come off quickly, and, as the bottle fills and the gas escapes into the air, dense white fumes will appear. This is because of the great affinity of this gas for

water. It condenses it from the air forming little droplets of hydrochloric acid solution. When the bottle begins to fume around the mouth, remove the delivery tube and hold it just above about a half-inch depth

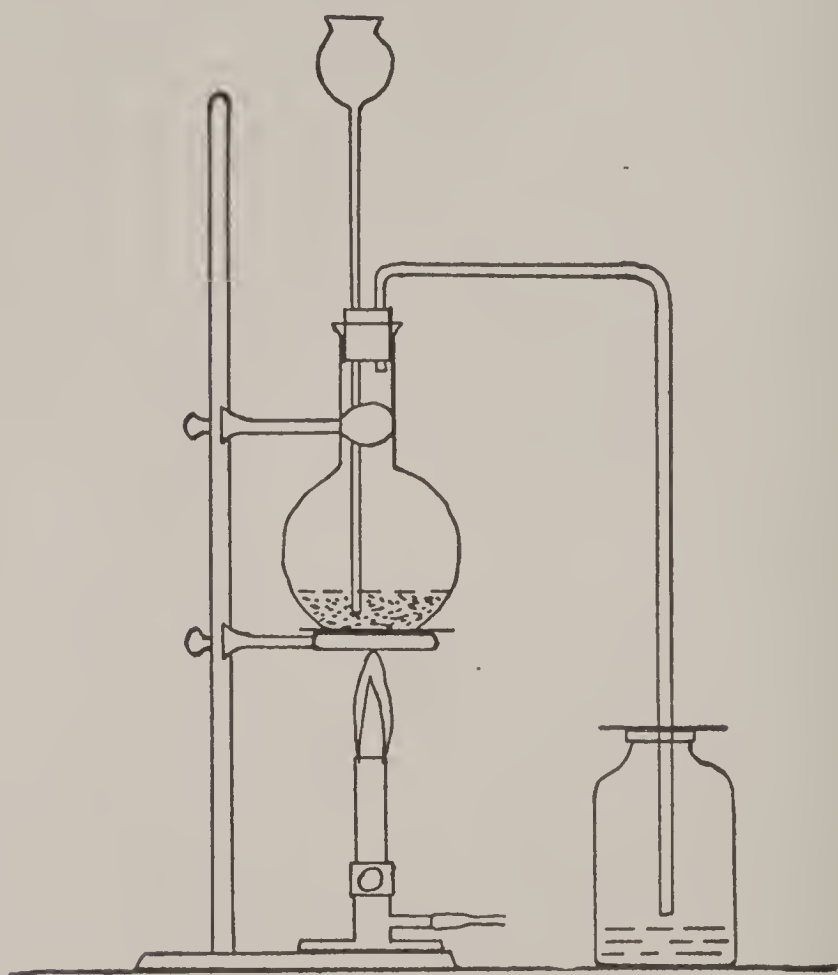


FIGURE 10.

Apparatus for the preparation of hydrochloric acid.

of water in the bottom of another bottle. Looking through the solution toward the light, note the oily appearance of the water, as the gas dissolves.

Quickly invert the bottle of gas, which you have prepared, in a basin of water colored with blue litmus. If the bottle is full of the hydrochloric acid gas, the

water will rise into it with a rush and at the same time turn the blue litmus red. Hydrochloric acid is the second most soluble gas known.

Pour a little strong ammonia water into a beaker and bring near to it the delivery tube of the generator. If the hydrochloric acid is still coming, dense white fumes will form. The hydrochloric acid and ammonia gases unite to form ammonium chloride, or sal ammoniac.

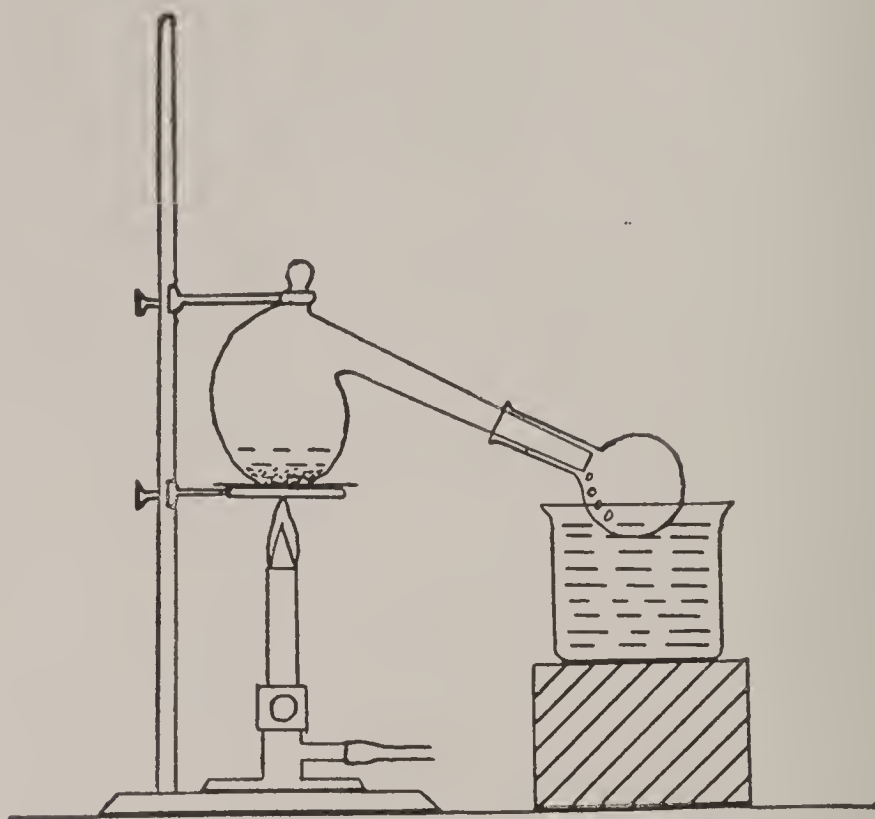
Test a little of the acid solution that you have prepared with litmus. Pour a little of it onto some magnesium ribbon in a test tube. Note the rapid action, and bring the mouth of the test tube near the flame. The combustion which you obtain shows the presence of hydrogen. Repeat using zinc or iron nails in place of magnesium.

You have now established two of the fundamental properties of an acid. It changes blue litmus red, and liberates hydrogen with a metal.

Preparation of Nitric Acid.—Nitric acid is a very active and interesting compound. It is one of the most powerful oxidizing agents known to chemistry and of great importance in the manufacture of explosives. Unlike hydrochloric acid, it is a liquid and may be distilled from a mixture of saltpeter and sulfuric acid.

Place in your glass-stoppered retort (see Forms of Apparatus) two tablespoonfuls of either sodium or potassium nitrate. These are salts of nitric acid, and just as we obtained hydrochloric acid from one of its

salts, so we shall employ the same method here. To introduce the salt into the retort place it upon a square of smooth paper and holding the retort in the left hand and the paper in the right slide the salt in. Then through a funnel pour on the nitrate about 15 c.c. of



Distilling nitric acid from a mixture of sulfuric acid and Chili saltpeter.

FIGURE 11.

strong sulfuric acid. Mount the retort on a ring-stand and thrust the neck into a flask resting in a jar of cold water (Fig. 11). Gently heat the retort and very soon the vapor of nitric acid will appear and begin to condense in the neck of the retort and run down into the flask. During the process rotate the flask so as to cool it more and make it a better condenser for the acid. Continue the action as long as acid distils

over. Then allow the contents of the retort to cool, add water and let it stand until the salt cake left in the retort has dissolved. But do not add water until the contents of the retort have cooled.

Oxidizing Action of Nitric Acid.—Into a test tube pour 2 c.c. of the strong nitric acid which you have just prepared. Stuff into the tube just above the acid a plug of excelsior and heat both the acid and the excelsior in the flame. A very brilliant combustion will occur. The nitric acid decomposes liberating oxygen, which unites with the combustible excelsior.

In your evaporating dish heat some sawdust until it just begins to char. Then pour upon it 3 c.c. of the nitric acid. A very vigorous combustion is set up.

A mixture of sugar and nitric acid warmed in a test tube produces a similar result.

On the wire gauze of your tripod or ring-stand place a small beaker or Erlenmeyer flask containing a mixture of sodium nitrate and sulfuric acid. Heat this gently and by means of tongs thrust into the hot vapor of nitric acid, which soon escapes, a glowing splint. The splint will instantly ignite and burn brilliantly. As it does so, brown fumes of nitrogen peroxide will also appear.

These experiments show the great oxidizing power of nitric acid and how unstable a compound it is. It is for this reason that this acid is used in the manufacture of high explosives.

Action of Nitric Acid on a Metal.—Pour 2 c.c. of the acid upon a copper rivet in a test tube. Add a

drop or two of water and a most vigorous action will result, accompanied by dense brown fumes of nitrogen peroxide. Nitric acid, unlike sulfuric and hydrochloric, does not liberate hydrogen when acted upon by a metal. This is because the hydrogen is oxidized by the acid as fast as it forms.

Preparation of a Base.—To a little ammonium chloride (sal ammoniac) in a test tube add a few c.c. of sodium hydroxide solution and heat. Hold in the gas a piece of moist red litmus paper. It will be turned blue showing the presence of a base. Smell of the gas cautiously. What is it?

Prepare a mixture of equal parts of sal ammoniac and dry slaked lime. Place it in a test tube fitted with a stopper and delivery tube exactly as you did in the preparation of oxygen. Clamp the tube to a ring-stand and thrust the delivery tube upward into an inverted bottle. The delivery tube should pass through a cardboard which acts as a cover for the bottle. Now heat the test tube gently along the whole length of the mixture, and the bottle will quickly fill with ammonia gas. Thrust the inverted bottle in a basin of water and note how rapidly the gas dissolves. If the water has been colored red with litmus to which a drop of sulfuric acid has been added, as the water rises in the inverted bottle, it will be turned blue by the action of the ammonia, which is a base.

Sodium and potassium hydroxides, the two strongest bases, may be prepared directly by the action of the metals upon water. The action in each case is vigorous

with the liberation of hydrogen. In the case of potassium the hydrogen takes fire giving a lavender to violet colored flame.

Sulfuric acid is one of the three most important raw materials of chemical manufacture. It is produced in enormous quantities. There are two commercial processes, but each depends upon the oxidation of sulfur dioxide and the absorption of the product in water.

You may prepare this acid by passing sulfur dioxide into hydrogen peroxide. In your hydrogen generator place 3 teaspoonfuls of sodium sulfite. Pass the delivery tube to the bottom of a test tube half full of hydrogen peroxide. Through the thistle tube of the generator pour hydrochloric acid. The sulfur dioxide which is generated will unite with the hydrogen peroxide to form sulfuric acid. Note that the test tube gets quite warm from the heat liberated in the reaction. Divide the solution into two portions. Test one with litmus. A decided acid reaction will be obtained. To the other portion add barium chloride solution and you will obtain a white precipitate which will not dissolve in hydrochloric acid. These tests prove that you have prepared sulfuric acid.

Testing Soils.—Bring moist samples of various kinds of soil in contact with moist red and blue litmus paper. After a time examine them. Old soil will give an alkaline reaction. When you have found a soil that does this, mix a basin full of it with 20 to 25 g. of well pulverized burned lime. Stir this thoroughly

and test again. Do you now understand why farmers use large quantities of lime on their plowed fields? Into another basin full of acid soil stir a handful of wood ashes. Again make the litmus test. Wood ashes contain potassium carbonate, a substance which like a base will neutralize an acid. In addition the element potassium is one of the three essential constituents of commercial fertilizers.

CHAPTER II

THE EXAMINATION OF WATER

WATER will dissolve more substances than any other one liquid. Its purity is absolutely essential to the health of people everywhere. It touches life in countless ways. We have seen that it is composed of hydrogen and oxygen. Chemically, it is a most interesting substance, and its study will afford us a great deal of real laboratory work.

Natural water is never pure. In the chemical sense we mean by pure water, water free from organic and mineral matter. Water may be fit to drink and yet contain mineral matter in solution. Since water is the best solvent known, it is all the while taking into solution foreign matter. The purest water in nature is rain water, but even this contains gases washed from the atmosphere and it is discolored from running over dirty roofs. You do not need to be told that the water we drink is of the utmost importance. In many industries, too, water must be selected with the greatest care. Iron in water is objectionable for dyeing, tanning, paper making, bleacheries, and laundries. Chlorides are harmful in the refining of sugar and the tanning of leather. Hard water cannot be used in laundries without softening, and it is the chief cause of boiler

scale. For drinking, water must be free from harmful bacteria and organic substances and not contain too much mineral matter. I think you already see that this examination of water is going to open up a fascinating field for the amateur chemist.

First let us look into the matter of purifying water. There are two principal methods—filtration and dis-

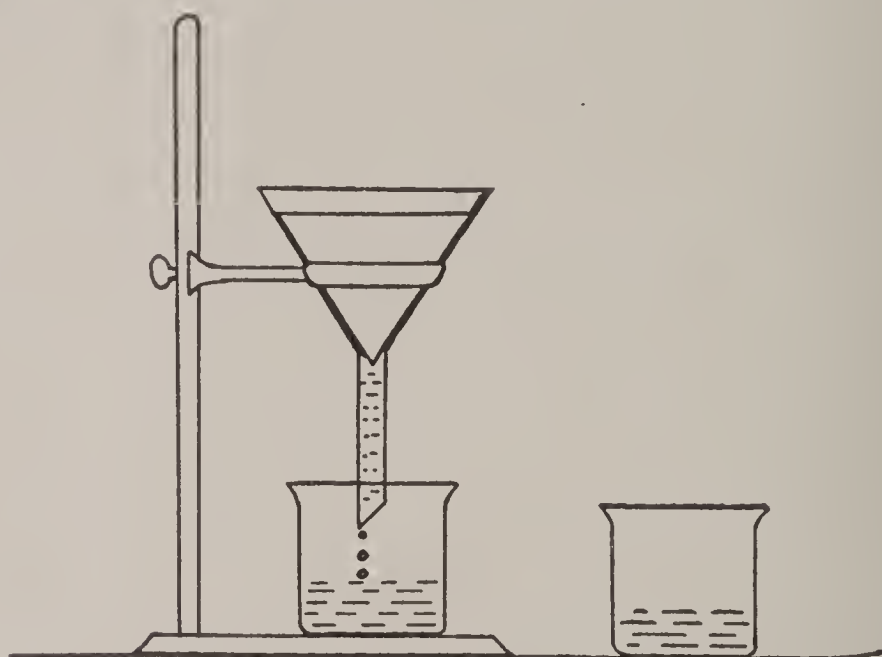


FIGURE 12.

Simple filtration for separating a solid suspended in a liquid.

tillation. Filtration removes sediment and may remove organic matter. Distillation removes all soluble matter and of course destroys any bacteria that may be present.

Filtration.—Stir a little soil into a beaker of water. Then fold a filter paper and fit it in your funnel placed on the ring-stand support. Pour the muddy water through the filter and catch the clear liquid that runs



Distillation of water.

through, called the *filtrate*, in another clean beaker (Fig. 12). This is the method always employed in the laboratory for the removal of suspended sediment.

Prepare a solution of ordinary table salt and pass it through another filter. Taste of the filtrate. This sort of filtration, you see, does not remove dissolved mineral matter. Confirm this by filtering a solution of blue vitriol, but do not taste it.

Filtering through Boneblack.—In a small beaker place 25 c.c. of cider vinegar and stir into it two teaspoonfuls of boneblack, or animal charcoal. Heat the mixture nearly to boiling for a few moments. Then pour it through a fresh filter into a clean beaker. Note that the vinegar has been decolorized.

Repeat this operation with a fairly strong solution of blue vitriol. This time the color remains. You have demonstrated a fact of considerable commercial importance, that boneblack will remove organic but not mineral coloring matter. In the refining of sugar boneblack filters are used to decolorize the syrup.

Another Method for Removing Color.—Natural water is frequently highly colored from having passed through peat beds or masses of decaying vegetation. In the manufacture of paints and pigments, organic dyestuffs must be removed from solution. A gelatinous substance known as aluminium hydroxide is very extensively employed in these decolorization processes. To learn the action of such a filter, prepare a dilute solution of ordinary alum by dissolving 5 g. of the salt in 250 c.c. of water. Also have at hand limewater,

made by shaking slaked lime with water in a stoppered bottle and allowing the mixture to stand for several hours. Then pour off the clear liquid. To one-sixth of a test tube of the alum solution add twice that volume of limewater and warm gently over the flame. Note the white gelatinous precipitate that forms. *A precipitate* is an insoluble substance which is formed when two solutions are mixed. If ammonia water is substituted for the limewater, a heavier precipitate may be obtained.

To show how this precipitate may be used to remove coloring matter prepare dilute solutions of some organic dyestuff. Logwood or alizarin are good for this purpose. Then to half a test tube of the alum solution add a little ammonia water, just enough to form a thin precipitate. Follow this with enough of the dyestuff to give a distinct color. Shake the mixture well and set the tube away. Within a short time you will note that the precipitate is settling and carrying down with it the coloring matter. After a time you will find a clear liquid above and a sediment of color-substance at the bottom. This color-substance is called a "lake," and such lakes are employed very extensively in the preparation of pigments. This gelatinous precipitate, known chemically as aluminum hydroxide, is also used to absorb the coloring matter from water.

A Sand Filter.—You may illustrate the action of aluminum hydroxide on water and also the efficiency of sand and rock in filtering natural water by making

a sand filter. Obtain a bottle of either a half-gallon or gallon capacity. Cut off the bottom according to the directions already given in the chapter on hydrogen and oxygen. Mount this, neck down, either on your ring-stand or on a support made by cutting a hole through the top of a box and knocking off one side. Fit into the neck of the bottle a rubber stopper carry-

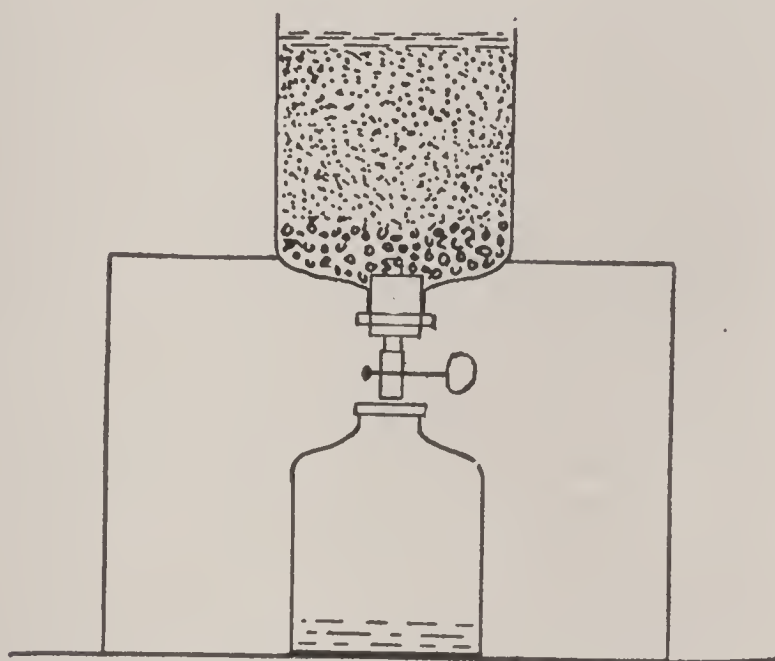


FIGURE 13.

A sand filter for the purification of water.

ing a short length of glass tubing to which you have attached a piece of rubber tubing and a pinch-clamp. In the bottom of this container place some clean gravel. Follow this with an inch layer of well washed sand. Then sprinkle in a very thin layer of powdered alum and slaked lime. Add alternately layers of sand and small amounts of alum and lime, until the bottle is nearly full (Fig. 13). Pour water through the filter

and, if it is not at first clear, continue to do so until the filtrate is free from sediment.

Now prepare several quarts of water colored with a small quantity of some dyestuff such as alizarin or magenta. The dyestuff may be dissolved in a little denatured alcohol, and only a few cubic centimeters of it will be required to give a deep color to a gallon of water. Pour this solution through your filter and draw it off by opening the pinch-clamp at the bottom. You will find that the water has been entirely freed from coloring matter. The alum and lime slowly unite to form the precipitate of aluminum hydroxide, and this together with the sand removes the organic matter. Such mechanical filters are used very much in the industries, and alum and lime are placed in city reservoirs to remove both sediment and organic coloring matter.

If possible, obtain a gallon of natural water containing both sediment and coloring matter and try the action of your filter upon it. Such a filter also removes bacteria. Do you not see why water from springs and deep artesian wells, which has seeped through many layers of sand and gravel, is pure?

Distilled Water.—For many purposes water must be free from both organic and mineral matter. Filtered water is never free from dissolved mineral matter. But for all work in exact chemical analysis, for storage batteries and the manufacture of ice this, too, must be removed. The only way to get this matter out is by distillation. In this process we boil the water

and condense the steam. Nothing but the steam passes over. The mineral matter is left behind.

A simple still may be made from a flask, a one-hole stopper, a bottle, a test tube, and a piece of bent tubing. (See Fig. 14.) The test tube surrounded by the

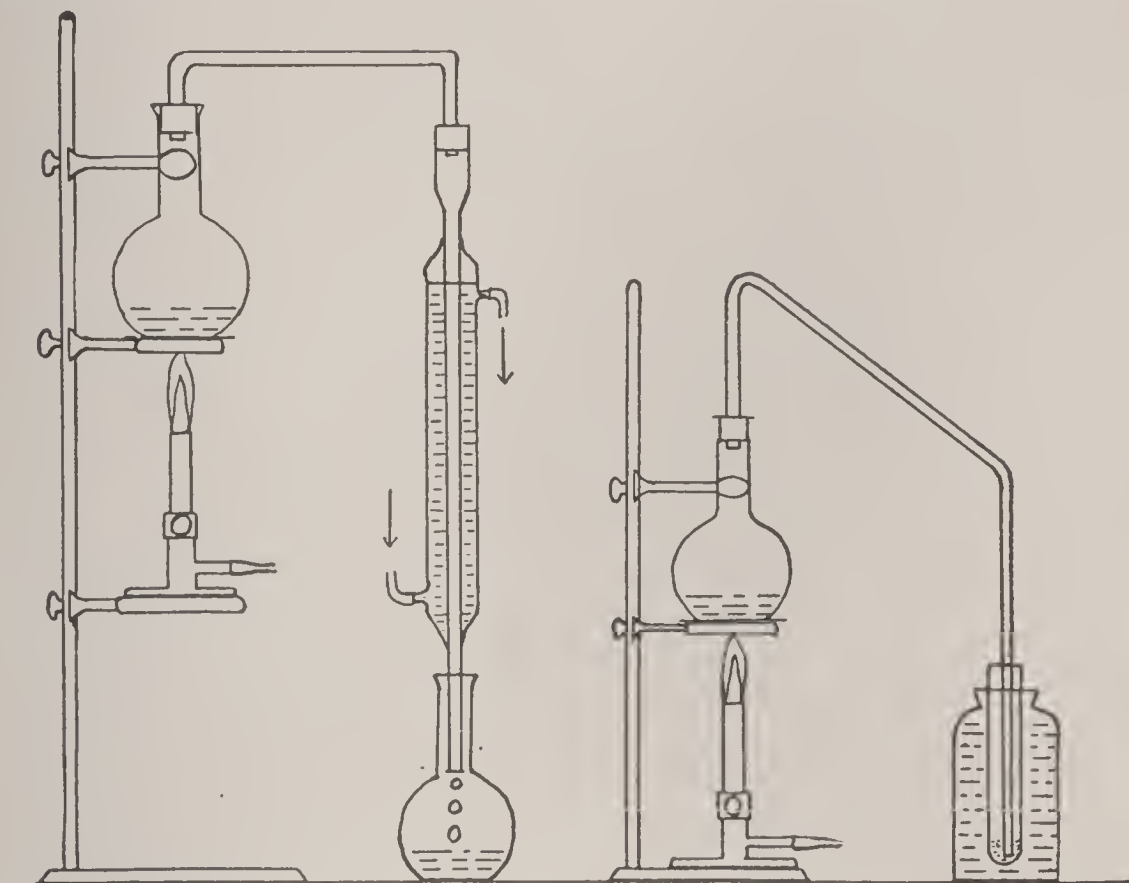


FIGURE 14.

Two set-ups for the distillation of water.

cold water in the bottle will condense the steam. The cooling water must be changed frequently and the flame should not be large.

When the still is ready, place in the distilling flask a solution of blue vitriol and light the burner. Let the water boil gently and soon you will notice the condensation of steam in the cooled test tube. This water that

distils over is called the "distillate," and you will note that it is colorless. It is pure water. It contains neither mineral nor organic matter. Place a little of the water in your mouth. You may think that it tastes "flat," but it is really tasteless. Pure water has no taste. This is the kind of water with which chemists do their work.

Empty out the distilling flask and substitute for the blue vitriol a dilute solution of some dyestuff. As this distils over note that the liquid, which you collect, is colorless and tasteless. The dyestuff is an organic substance.

To show how natural water may be purified, from a stagnant pool obtain a sample containing both sediment and coloring matter. Filter it through the sand filter. It will come through free from sediment and color. But place a little of the filtrate on a watch glass and evaporate it over steam. As I have already told you, to do this place the watch glass on the top of a half-filled beaker of water and heat with the Bunsen flame. If there is no lip on the beaker, draw the watch glass slightly to one side so as to allow the steam to escape. As the water evaporates, you will observe a white residue on the watch glass. This is mineral matter which the sand filter could not remove.

Now distil some of the water that passed through the filter and again evaporate a little of the distillate over steam. This time you find no residue.

Your distillation experiments can be carried out on a much larger scale if you have a Liebig condenser,

as shown at the left in Fig. 14. To use it you will need running water and a sink. If your laboratory does not contain them, possibly you could use the kitchen.

The chemical purification of water for drinking purposes is also important. To destroy bacteria and organic matter the poisonous gas *chlorine* and the powerful oxidizing agent *ozone* are employed, besides aeration. In the latter process the water is thrown upward in a spray and mixed with the air which oxidizes the organic matter. These methods, you will be unable to demonstrate.

Sanitary Examination of Water.—If you are to be a real chemist you must be able to determine whether water is pure and wholesome to drink. Has it suffered contamination from sewage, sinks, or products or organic decay? Is it free from an excessive quantity of mineral matter? Does it contain lead dissolved from the pipes through which it has passed? These and other questions you must be able to answer.

Taking the sample is of the utmost importance. In doing so note the surroundings. Discover whether there is any possible source of contamination. Then be sure that the bottle in which you take the sample is *chemically* clean. Simply washing with soap and water is not sufficient. Fill it a quarter full of strong sulfuric acid to which you have added a small quantity of potassium bichromate crystals. Let the solution stand for several hours, shaking at frequent intervals. Then pour the solution into a glass-stoppered bottle and preserve for future use. Rinse the bottle well. Fill

it with the water to be examined. Throw this out and fill it again. Take the sample to the laboratory and begin its analysis at once.

The amount of *sediment* in water, its *color*, and its *odor* are indications of its purity. Allow a test tube of the water to stand overnight, unless it is perfectly clear. Pour off the liquid and if there is a residue examine it with a hand magnifier. If bits of hair or cotton and woolen fibers are present, they might indicate organic filth. To test the color, fill a long test tube with the water and stand it on a piece of white paper in front of a window giving a good light. Back of the tube place a piece of white paper extending almost to the bottom. If on looking down into the tube, the water is transparent, or shows only a slight bluish tinge, the color is natural. If, however, there are tints of brown, yellow or green, the water may be polluted. To determine whether the water has odor, fill a flask half full and stopper it. Heat it until the glass feels thoroughly warm. Shake it, unstopper, and smell of the gas that has been driven from the water. Pure water will give no odor. If the odor is putrid, the presence of decaying organic matter is indicated, and it may be either of animal or vegetable origin.

The *total solids* in water is another important item. By total solids is meant the residue of mineral and organic matter left upon evaporating water to dryness. To make a real determination of this, you must have a balance. A hornpan balance will do. Weigh in it your evaporating dish to the nearest milligram,



that is, 0.001 g. Your set of weights should permit you to do this. Place in it 100 c.c. of the water, or half that amount if your dish is small. The water may be measured either with a graduate or a pipette. Evaporate this over a half-filled beaker of water to complete dryness. Then heat it over a very small flame for 30 minutes. When the dish has cooled so you can handle it, weigh it again. The increase in weight gives the total solids. If 100 c.c. were taken, the increase in weight in milligrams multiplied by 10 will give you the number of parts per million of solids in the water. For drinking purposes this should not be over 600.

The next step is to determine the presence of *organic matter*. To do this evaporate to dryness, as above, 50 c.c. of the sample in a porcelain dish. When dry, place the dish over a Bunsen burner and heat very gradually to a considerable temperature. If the residue chars and blackens, organic matter is present and at once indicates contamination. The organic matter in itself is not harmful, but it may be associated with disease-producing bacteria.

Chlorine, in the free state, is a deadly poison. In the form of sodium chloride it is essential to life. Although chlorine in combined form is not harmful, natural water contains it in only very small quantities. If much is present, it usually means sewage contamination.

The *qualitative test* may be very simply made. Dissolve a small crystal of silver nitrate in 10 c.c. of dis-

tilled water. (A druggist will supply you with distilled water, or you may prepare it if you have a Liebig condenser.) Add a pinch of table salt to a half test tube full of tap water and then a few drops of the silver nitrate solution. A white precipitate of silver chloride will be obtained. This is the test for a chloride. Make a very dilute solution of the sodium chloride, not over a half dozen grains to a test tube of water, and repeat the test. This time you obtain only a cloudiness in the water, and yet it is probably deeper than what you will get with any natural water that is fit to drink.

Now make the test on the sample of water which you are examining. Pure water should contain very little chloride. If the test gives more than a faint cloudiness, the water should be regarded with suspicion. There may be some perfectly harmless source of the chloride, such as a field fertilized with chloride of potash, and that is where a knowledge of the conditions surrounding the water supply is of the greatest importance.

For those amateur chemists who have the equipment I shall describe the *quantitative test for chlorine*. For this determination you will need a *standard solution of silver nitrate*. This is a solution prepared by dissolving exactly 2.394 g. of silver nitrate crystals in a liter of distilled water. A liter is 1000 c.c. Half or even a quarter of that volume will be sufficient. Suppose we prepare 250 c.c. by dissolving 0.599 g. of the silver salt in that volume of water. Mount on

your ring-stand support a burette and nearly fill it to the top graduation mark with the solution.

To read a burette place your eye on a level with the surface of the liquid in it and take the reading from the lowest point of the curved surface, estimating the number to the nearest tenth of a c.c. The difference between two successive readings will be the number of c.c. used.

This will be another titration experiment and you will need an indicator. Prepare it by dissolving 5 g. of potassium chromate in 10 c.c. of water. To learn the effect of silver nitrate upon this, add 3 or 4 drops of the indicator to a little water in a beaker and run in from your burette a few drops of the standard solution. A red precipitate will form. Then prepare a very dilute solution of sodium chloride and add to it the indicator. As you run in the silver nitrate this time, note that white silver chloride at first forms, after which with continued addition of the standard solution the red precipitate appears. But this will not happen until all of the chloride has been precipitated.

Now to make the real test, transfer to a small beaker 50 c.c. of the water to be examined and add to it 3 or 4 drops of the indicator. Place this on a white paper under the burette near a window giving good light. Record the reading of the burette and begin to add the standard silver nitrate solution a drop at a time, stirring after each addition with a clean glass rod, which must not be removed from the beaker. Continue this until a faint tinge of red appears. When

this happens you may know that the chloride has just all of it been precipitated. Take the final reading of the burette and subtract the first reading from it. Each cubic centimeter of silver nitrate used means 10 parts of chlorine per million parts of water. In the most exceptional cases there should never be more than 50 parts per million, and usually the amount is less than 10.

Nitrogen.—Another highly important indication of the fitness of water for drinking purposes is the *presence of nitrogen compounds*. They may be found as ammonia, nitrites, or nitrates. Whenever nitrogen in any form is found in water, it means contamination from organic sources. The nitrogen itself is not harmful but with it will usually be found germs of disease.

The tests for these substances are very delicate and they cannot be carried out in a room containing any ammonia bottles.

To test for ammonia a solution known as *Nessler's Reagent* will be required. Get your druggist to prepare a small bottle of it for you. It would be rather difficult for you to prepare it yourself. Place 50 c.c. of the water to be tested in a tall test tube and add to it from a pipette or a burette 2 c.c. of the reagent. After 5 minutes look down into the tube. A yellowish brown tinge proves the presence of ammonia.

But ammonia may also be present in combined form and in that case the above test would not show it. To test for ammonia compounds prepare a solution consisting of 125 c.c. of distilled water, 25 g. of potas-

sium hydroxide, and 1 g. of potassium permanganate. Fill your distilling flask half full of the water to be tested and add to it 25 c.c. of this solution. Into tall test tubes distil off successive portions of 50 c.c. each. Throw the first portion away, for the apparatus itself might have contained traces of ammonia compounds. To each of the other portions add 2 c.c. of Nessler's Reagent and at the end of 5 minutes examine the color. A yellowish tinge tells the story.

To detect *nitrites* two solutions will be required. *Sulfanilic Acid* solution is prepared by dissolving 1 g. of the acid in 100 c.c. of hot water. To prepare *naphthylamine hydrochloride*, the other reagent, boil a half gram of the salt in 100 c.c. of water for 10 minutes.

To make the test place 50 c.c. of the water in a tall test tube and add 1 c.c. of concentrated hydrochloric acid, that is, the strongest acid. Follow this with 2 c.c. each of sulfanilic acid and naphthylamine hydrochloride. Cover the test tube with a glass plate and allow it to stand for 30 minutes. If nitrites are present, a pink color will appear. By the intensity of the color you may estimate the amount of nitrite. Wholesome water seldom contains even the faintest trace.

To test for *nitrates* get your druggist to prepare for you a solution of *phenol-sulfonic acid*. Evaporate 100 c.c. of water to dryness in a porcelain dish over steam and add to it 2 c.c. of the above reagent. Wholesome water will give no red color within 10 minutes.

At the end of that time add a few c.c. of ammonia water. If nitrates are present, a yellow color will appear. Since sodium nitrate is a common fertilizer, nitrates are more apt to be present than nitrites, but in no case should there be more than a trace.

You may also test for nitrates by adding to 50 c.c. of the water 2 c.c. of concentrated sulfuric acid and a minute quantity of a compound known as *brucin*. A red coloration proves their presence.

Sometimes water contains *lead* dissolved from the pipes through which it has passed. To test for it evaporate 100 c.c. of the water to one-tenth that volume and add a few drops of potassium chromatic solution. If lead is present, a yellow precipitate will form.

Good drinking water will never contain more than traces of *phosphates*. Again get your druggist to prepare for you a solution of ammonium molybdate. Then evaporate 100 c.c. of the water to dryness and moisten the residue with a few drops of the reagent. Warm gently and look for a yellow color. If it appears, phosphates are present, and this would indicate animal contamination.

Another method of testing for organic matter is by the oxidizing action of potassium permanganate. Prepare a solution of this salt by dissolving 0.2 g. in a half liter of water (500 c.c.). Also very slowly pour 20 c.c. of concentrated sulfuric acid into 175 c.c. of distilled water and allow the mixture to cool.

Into a clean porcelain evaporating dish measure 100 c.c. of the water and add 10 c.c. of the acid. Heat the



Testing for organic matter in water.

contents of the dish nearly to boiling and add from your burette, drop by drop, the solution of potassium permanganate, stirring and heating after each addition. Continue this process until a permanent tinge of pink appears. If organic matter is present, it will take oxygen from the potassium permanganate and decolorize it. If more than 4 or 5 c.c. of the permanganate are decolorized by 100 c.c. of water, there is probably some source of pollution.

These water tests are giving you a great deal of practical laboratory work, and I am sure that you are enjoying them. There is never anything much more fascinating than doing real analytical work.

Mineral Analysis of Water.—Although to know whether water is pure and fit to drink is the most vital question concerning it, still for industrial purposes the mineral compounds present must be known. Probably the most important item that comes under this head is that of *hardness*. You know that some water feels harsh to the touch and roughens and chaps the skin. Other water like rain water feels soft and has just the opposite effect upon the skin. These were the original meanings of the terms “hard” and “soft” as applied to water. But the action of water with soap is now taken as an indication of its hardness.

To test for the degree of hardness obtain a good quality of liquid soap as free from color as possible. Place some of this in a test tube fitted with a cork in the opposite sides of which you have cut two small notches. This will serve as a dropping bottle.

Distilled water is perfectly soft. To two-thirds of a test tube of this water add one or two drops of soap. Observe that the soap produces no cloudiness in the water. That is the first indication that water is soft. Next place your thumb over the mouth of the test tube and shake it thoroughly. Copious suds at once form and do not disappear for several hours.

Now add a few drops of a solution of calcium chloride to a half test tube of water and follow it with soap. Note that cloudiness appears in the water and a precipitate forms. Shake the tube. Froth and bubbles may form, but not suds. This is typically hard water. Natural water is not usually as hard as this prepared sample but the action is the same. The precipitate that forms is the cause of the scum which appears on the sides of a wash bowl or bath tub. In the laundry it is difficult to wash this substance from the clothes and it behaves just as so much dirt. In fact dirt is defined as matter out of place.

Prepare another sample of hard water by dissolving a pinch of magnesium sulfate, ordinary Epsom salt, in half a test tube of water. Test this with soap. You will observe that a precipitate forms again and that no suds appear on shaking. Calcium and magnesium compounds are the causes of hard water in nature. But there are two kinds of hardness—*temporary* and *permanent*. One may be removed by boiling while the other requires chemical treatment.

Temporary hard water in nature is produced by the solvent action of water containing carbon dioxide in

passing over limestone rocks. It is this action that hollows out limestone caverns and builds stalactites and stalagmites. You may prepare temporary hard water by passing carbon dioxide through limewater until the precipitate which first forms dissolves and the water becomes clear.

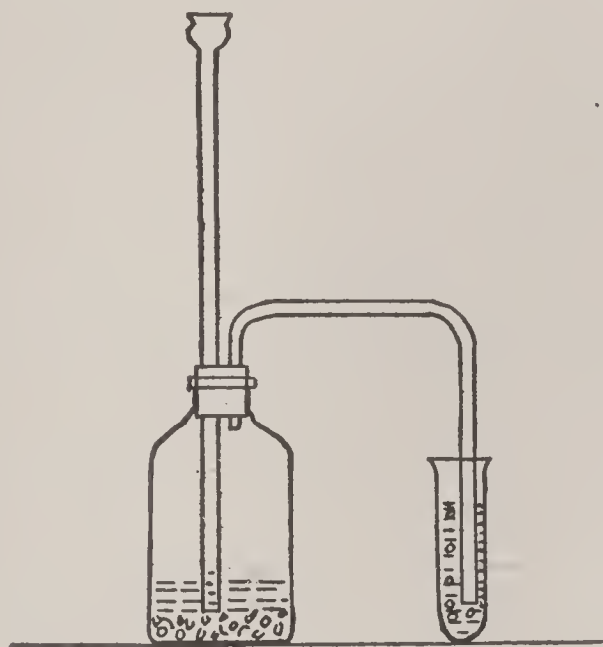


FIGURE 15.

Preparation of temporarily hard water by passing carbon dioxide through limewater.

Set up your gas generator as you did in the chapter on the atmosphere and with marble chips and hydrochloric acid prepare the gas. Pass the delivery tube to the bottom of a test tube nearly full of limewater and bubble the gas through it until the white precipitate that first forms entirely dissolves (Fig. 15). Divide this solution into two portions. Dilute one with a little water and add soap. You find it hard of course. Heat the other portion in the flame. The

white precipitate that dissolved in an excess of carbon dioxide will reappear. When this happens, filter off the precipitate, dilute the filtrate with a little distilled water, and add soap. This time you will find the water soft. Temporary hard water is one of the chief causes of the "boiler scale" that forms on the inside of tea kettles, steam boilers, and hot water pipes. Place some of the scale from the inside of a tea kettle in a test tube and add hydrochloride acid. You will obtain a vigorous effervescence of gas, which you may test with limewater for carbon dioxide.

Prepare another sample of the temporary hard water and add to it limewater. You will observe the formation of a white precipitate. This is another method employed to soften such water for industrial use.

Permanent hardness is due to the presence in water of calcium and magnesium chlorides and sulfates. Shake a little Plaster of Paris (calcium sulfate) with a test tube of water, filter, and divide the filtrate into two parts. To one portion add a fairly strong solution of ordinary washing soda, or sodium carbonate. A white precipitate will form. Filter this off and test the filtrate with soap. Copious suds will form and you will find that the water has become quite soft. This is the most common method of softening permanently hard water. In the same way the hardness may be removed from water containing magnesium sulfate.

Boil the other portion of the above filtrate and note

that a white precipitate forms. This is because calcium sulfate is less soluble in hot water than in cold. This is one of the causes of boiler scale. Filter off the precipitate and test the filtrate for hardness. You will find it still hard. And, because such hardness cannot be removed by boiling, it is called permanent.

A large scale method of removing hardness has recently come into very wide use in laundries, textile mills, and wherever industries require soft water in large quantities. It consists in passing water through immense filters containing an artificially prepared mineral substance, known under a variety of trade names such as "permutit," "softite," "refinite," etc. The water issues from these filters perfectly soft. After a certain period of use, the filter must be regenerated, but this is easily done by allowing a strong solution of brine to stand in it overnight. Upon drawing off the brine and washing out the filter, it is good for another period of service.

Now to test the sample of water which you have for hardness add to 25 c.c. of it soap, a drop at a time, shaking the test tube after each addition, and count the number of drops required to give permanent suds, that is, a small depth of suds which do not disappear immediately after shaking. To estimate the degree of hardness compare the amount of soap used with the water being tested and with that required for an equal volume of distilled water.

To determine whether the hardness is temporary or

permanent boil 25 c.c. of the water and test with soap. If no more soap is now required than with the same volume of distilled water, the hardness is all temporary. If the boiling seems to make no difference in the amount of soap required, the hardness is all permanent. If the amount of soap required is less than it was before boiling but more than with the same volume of distilled water, the hardness is of both kinds.

You should learn to test for a number of the important mineral constituents of water. These are usually present in quite small quantities and, in order to obtain good tests, it will be necessary to evaporate a quart of water to the volume of about a half a pint. This may be done in a granite basin over the kitchen stove. If a precipitate forms, filter it, catching the filtrate in a clean beaker. Replace the beaker with a test tube and pass through the filter paper a little hydrochloric acid. If in doing this effervescence occurs, *carbonates* are present.

To a portion of the main filtrate add a few drops of a solution of barium chloride. If a precipitate forms which will not dissolve in dilute hydrochloric acid, *sulfates* are present in the water.

If another portion gives with silver nitrate a white precipitate, which is insoluble in dilute nitrate acid, *chlorides* are present.

Iron may be tested for by adding to a little of the filtrate a few drops of potassium sulfocyanate solution or potassium ferrocyanide solution. A red color

in the one case or a blue color in the other shows the presence of iron.

Calcium salts may be tested for with a solution of ammonium oxalate. Add a few drops of the reagent to a little of the filtrate and note whether a white precipitate forms. If so, calcium is present.

The test for *lead* has already been given.

We have now given water a very thorough examination and you have gained by it a considerable knowledge of chemistry and excellent drill in laboratory work. But we have only begun to experiment. Some of the most interesting work lies ahead.

CHAPTER 12

SIR HUMPHRY DAVY

IF you should suddenly find yourself the master in the best equipped laboratory in the land surrounded with all the conveniences that money and science could provide, fortunate, indeed, would you consider yourself. But such was the happy fate that befell Sir Humphry Davy, when he was not Sir Humphry, but just plain Humphry Davy. In many respects young Davy lived in a most favored time for amateur scientists. All scientists were more or less amateurs in those days, and it did not require a college training and a number of degrees to obtain real positions in the scientific world.

Davy was born in Penzance, Cornwall, England, in 1778, the son of a woodcarver. One of his early pastimes was to experiment in a home laboratory in the loft of the house. And more than once his chemical mishaps aroused the opposition of his elders, who feared that he would some day blow them all into the air. But experiment, Davy would, for he was a born scientist. Yet Davy was a genuine boy. He loved to fish and he did not do very well in school. At the death of his father, Davy was apprenticed to a neighboring physician and assisted him in the preparation

of his medicines. This was just what the young scientist wanted, for now he had a real laboratory in which to experiment. But the numerous explosions which he devised soon led his employer to let him go.

In the meantime Davy had fortunately secured an introduction to Dr. Beddoes of Bristol, who was just establishing a "pneumatic institute" for the preparation of new gases and the discovery of their physiological effects upon the human system. Dr. Beddoes was looking for a director of his institute and, learning of Davy's great interest in chemistry, asked him to accept the place. Davy was delighted and immediately took up his new duties. Imagine it, if you will, an untrained home-laboratory worker placed at the head of an important scientific institution. But he plunged into his new work with zest and quickly discovered the intoxicating and soothing effects of the anæsthetic called nitrous oxide, but better known as "laughing gas." It is the same gas that dentists use today, but Davy was the first to experiment with it. He discovered that upon breathing it the patient first becomes hysterical and then unconscious. News of the discovery quickly spread throughout England and Davy found himself famous. Everyone wanted to try the properties of the new gas, and demonstrations with it became the fad of the hour. In breathing carbon monoxide, however, Davy became ill and nearly lost his life. It is a good thing that not many gases were known in those days or he surely would have been a martyr to science.

When Davy had just passed twenty-one, Count Rumford established the Royal Institution in London, and so renowned had the youthful scientist become that he was called to be assistant professor of chemistry in the first important scientific laboratory ever founded in England. As Davy sat in the stagecoach which bore him from Bristol to London, he dreamed of future wealth and fame, and well were these dreams fulfilled, for he became one of the foremost scientists of Europe. So popular were his lectures that all London flocked to hear him, and for his great discoveries the king knighted him. And all this came to a mere youth who taught himself chemistry and learned the science of laboratory technique from his own instruction.

CHAPTER 13

SOAP

AFTER our study of water I think you would expect soap to be the next subject. Soap and water usually go together. Soap-making is one of the oldest chemical processes known. Does your mother ever make soap at home? Our grandmothers did. I remember the old ash-leach into which we put all the wood ashes from the fires, and through which each day we poured a pail of water. The water leached down through and produced *lye*, which we caught in an iron kettle. This lye was a yellowish liquid and exceedingly caustic. Then every few weeks we placed all of the scraps of fat that had accumulated about the household into another huge iron kettle, called a caldron kettle, and raised about a foot off the ground by large stone blockings. Over these fats we poured the lye mixed with water. Next we built a fire and heated the contents of the kettle for several hours, stirring the mixture frequently with a wooden paddle. During this process the lye united with the fat to form soap. The name given to this soap-making process is *saponification*. When all the fat had been saponified the soap would slip freely from the wooden paddle and the process was at an end. This soap was always soft

and was kept in earthen jars. It was the only soap used in the early days, but it was very harsh on the skin and injurious to fabrics, especially woolens.

The preparation of soap even today employs the same method. We will start you off with a simple laboratory process and follow it with one on a larger scale. Pour into a 500 c.c. flask 50 c.c. of cottonseed oil and 15 c.c. of a solution of sodium hydroxide having 40 g. of the solid hydroxide to 100 c.c. of water. Fit into the neck of the flask a one-hole rubber stopper carrying a 2-foot length of glass tubing. Place the flask on your ring-stand and heat the mixture gently with a small flame for at least two hours. If 100 c.c. of alcohol have been added to the mixture the time of heating may be cut in half. Unless the alcohol is used, the stopper and glass tubing will not be required. At the end of this time pour the mixture into an enameled pan and boil it with constant stirring until all the alcohol has been driven off or until it becomes a pasty mass. Then let it cool.

Test a little of the product with soft water and note whether it produces suds freely. Wash your hands with it.

To prepare soap on a larger scale obtain from the grocery store a small can of concentrated lye and a pound can of Crisco. Place a half pint of cold water in a large beaker and dissolve in it 4 big tablespoonfuls of the lye, being careful not to get any of it on your skin or clothes. Put the Crisco into a 2-quart enameled pan or kettle and heat the fat until it melts



Making soap.

and becomes lukewarm. Then pour the solution of lye into the melted fat very slowly and with constant stirring. Continue the stirring until you obtain a thick, pasty mass. Pour the product into a shallow oblong basin and let it stand for a week. Cut the soap into small cakes and distribute them to your friends as evidence that you have become a manufacturing chemist.

You will find it very interesting to prepare special soaps. While the soap is still thin stir into it some dyestuff and you will have a *colored* soap. A dash of nitrobenzene or toilet water will give a *perfumed* soap. A 5 per cent solution of carbolic acid stirred into the batch gives a *medicated* soap. Mentholum or camphor may be substituted for the carbolic acid. Fine sand gives *scouring* soap. Oatmeal or cornmeal will give a "*beauty soap*."

Soft soap is made by saponifying a fat with potassium hydroxide instead of sodium hydroxide. Proceed exactly as you did with the first soap made above, but in place of the 15 c.c. of sodium hydroxide use 21 c.c. of a solution of its twin brother, potassium hydroxide, containing 40 g. to 100 c.c. of water. After the boiling is complete, pour into the soap 200 c.c. of water and mix thoroughly. Put this up in small bottles and distribute it to your friends.

Good toilet soap should contain no *free alkali*. To a freshly cut piece of toilet soap add a few drops of an *alcohol* solution of the indicator phenolphthalein. If free alkali is present a pink color will appear. Free

alkali is harsh on the skin and in laundry soaps injures woollens and delicate fabrics.

Hard Water Wastes Soap.—To determine something about how great this waste is place 100 c.c. of distilled water in a glass-stoppered bottle and add to it from a burette liquid soap a drop at a time followed by shaking until a suds that will remain permanent for 5 minutes is obtained. From the readings of the burette note the amount of soap used. Then repeat the test using tap or well water. The difference in the amounts of soap used in the two cases will enable you to estimate the waste resulting from the use of hard water. Of course this will vary with the water used.

The cleansing action of soap was for a long time more or less of a mystery. We think we now know about this. To understand the action of soap with grease we must know what an *emulsion* is. Place 2 or 3 drops of olive oil in half a test tube of water and shake the mixture. Set the test tube aside. In a few moments you will observe that the oil rises to the surface of the water. The two will not stay mixed. Now add to the test tube a few drops of liquid soap and shake again. This time you obtain a permanent mixture. The oil will not rise to the surface even upon standing for a long time. You have emulsified the oil, and that is exactly what soap does with grease on your hands or on goods that are being washed. The soap emulsifies the grease and causes it to slip off.

Just plain dirt actually adheres, or sticks to soap,

which wraps itself about the dirt in a tiny film and carries it away.

Another factor in the cleansing action of soap is its alkaline reaction. Dissolve a little soap in water and test it with phenolphthalein indicator. A pink color will always result showing the presence of an alkali, or a base. This is formed by the reaction of the soap and water with each other. Since an alkali will attack grease and dirt, it also assists in the cleansing action.

I think you will find the examination of a soap powder interesting. Washing powders usually contain ground up soap, sodium carbonate or borax, and often some gritty substance such as fine sand or pumice stone. Weigh out about 10 g. of the powder. Transfer it to a beaker and cover it with 200 c.c. of hot water. Stir this thoroughly to dissolve all the soluble matter. Then filter and wash the insoluble matter on the filter with hot water a number of times, passing the water through the filter. Examine this residue. If you wish to determine the per cent, dry the filter on a watch glass placed over a water bath and weigh the residue. The difference between this weighing and the original weight of powder taken will be the amount of insoluble matter plus the weight of the filter paper. To obtain an accurate result it will be necessary to weigh a filter paper and subtract it from the apparent residue.

To test for the carbonate place a little of the powder in a test tube and add hydrochloric acid. If effervescence occurs, incline the test tube so that the

mouth is slightly above the horizontal and let the gas slide into another test tube containing about 2 c.c. of lime-water, but do not let any of the liquid pass into the limewater. Shake the test tube containing the lime-water and note that a white precipitate forms, showing the presence of carbon dioxide and consequently of a carbonate.

To test for borax dissolve a little of the powder in hot water, filter the solution, and add to the filtrate hydrochloric acid. Dip a piece of *tumeric paper* into this solution and dry it by holding the paper high over the flame. If borax is present, you will note that the paper has a tinge of red which turns black upon the addition of ammonia water, but that dilute hydrochloric acid will restore the color.

To learn the *action of borax in soap* dissolve a little borax in a test tube of water and try the action of the solution on red litmus paper. The blue color that appears shows that borax gives an alkaline reaction and is therefore a cleanser. In addition borax softens water. Prepare a sample of hard water by shaking a little Plaster of Paris or Epsom Salt with a test tube of water and filtering. Add to the filtrate a little powdered borax, shake well, heat nearly to boiling and filter again. Test the filtrate with soap. Do you find that it has been softened?

In the manufacture of soap a very *valuable by-product* is obtained. It is glycerin so much used in the making of explosives. In the soap which we have made the glycerin remains in the soap.

CHAPTER 14

EXAMINATION OF TEXTILE FIBERS

WHEN you buy a suit and the salesman tells you that it is all wool, you want to be sure that he is telling you the truth. You do not want it to be half cotton. But how are you going to know. You will have to turn chemist. From early times fabrics have been woven from four principal kinds of fibers—wool, cotton, silk, and linen. Two of these are of animal origin and two of vegetable. In addition we now have artificial silk.

The burning tests are the easiest to apply, so we will take them first. Obtain samples of each of the four chief kinds of fabrics and ravel out some of the threads. Ignite with a match each kind in succession and note in each case the character of the flame, the odor, and the ash. I think you will find that the *cotton* burns up quickly, giving a bright yellow flame and little odor, and leaving a fine white ash. If much of the cotton is burned you will probably detect the odor of burning wood. Cotton and wood are similar in composition. *Linen* fibers behave in the same way. But you will find that the *wool* fibers do not burn so readily. They give a flickering flame with a disagreeable odor like that of burnt hair, and leave a hard

black residue of unburned carbon. These properties are characteristic of animal fibers when burned. Now try the *silk*. You will see that it resembles wool, but it burns more quickly, the flame is blue, and the odor is not as noticeable. But there is still an unburned residue. Try the burning test on *artificial silk*. Which one of the above fibers does it resemble? How could

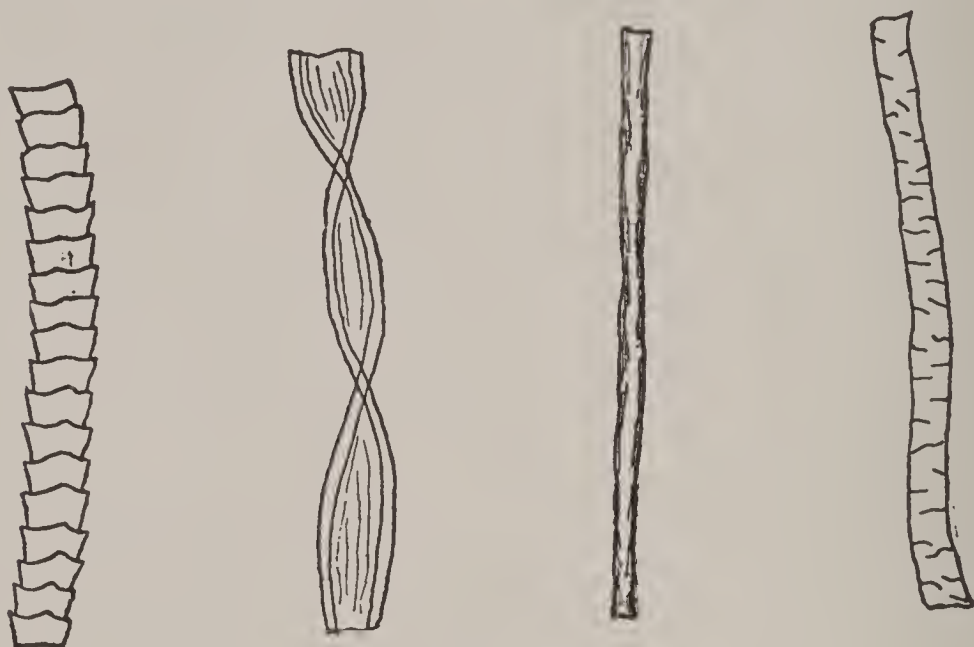


FIGURE 16.

Microscopic appearance of wool, cotton, silk, and linen fibers, named in order.

you distinguish real silk from artificial silk? I think you will find this easy to answer. How does *mercerized cotton* behave when burned? Would you have any difficulty in distinguishing it from either kind of silk?

A microscopic examination of fibers is often made. Possibly you will be able to detect some differences with a small hand magnifier, but, if by any chance you can get the use of a compound microscope,

do so. If you are a high school student, this can be easily arranged. Wool fibers seem to be made up of a succession of small segments, which overlap each other like the scales of a fish or shingles on a roof. A cotton fiber has the appearance of two twisted strands. Silk fibers seem to be long, narrow, cylindrical threads. Linen fibers are similar to silk but they are wider, not so straight, and often have cross lines.

Now for the chemical tests. The alkali test is best for distinguishing *cotton and wool*. From the grocery store obtain a small can of concentrated lye. Into a small agate basin placed over your Bunsen burner put 4 big tablespoonfuls of the lye. Add a pint of water and bring it to a boil. Then drop into this solution a square of some fabric, which you think is wool or partly wool and partly cotton. Boil it slowly for a few minutes. If the sample completely dissolves, it is all wool. If it does not, the insoluble residue is probably cotton. In case of any such residue, remove it with a glass rod and wash out the fibers thoroughly in several changes of cold water. Then dry them, and apply the burning test. Very often so-called "all wool" contains considerable cotton.

To distinguish wool and silk drop threads of each into small beakers containing concentrated hydrochloric acid. After a few minutes you will find that the silk has completely dissolved, but that the wool has been little affected. They are both soluble in lye. Of course you understand that, chemically, lye is sodium

hydroxide, and that either this or potassium hydroxide may be substituted for it.

To distinguish real silk from artificial silk or mercerized cotton place three small beakers upon the table, each containing concentrated hydrochloric acid. In the first put real silk, in the second artificial silk, and in the third mercerized cotton. Boil each for two or three minutes. Which dissolves? The burning test may also be used.

To distinguish cotton and linen is more difficult. There is no good chemical test. The best means is the microscope, the differences in the appearance of the fibers having already been described. Another test consists in boiling small pieces of the goods in olive oil for about five minutes. Upon examining them against a dark background, the linen looks dark, while the cotton appears light.

To determine the percentage of cotton in mixed goods cut a piece of the cloth about 3 inches square and dry it for several minutes in your evaporating dish placed high over a small flame. If you have a thermometer, keep the temperature at about 100 degrees Centigrade. Then allow the cloth to cool without removing from the dish. Transfer it to your balance and weigh it as accurately as possible with the set of weights which you have. Place it in a beaker and boil it with a solution of either sodium or potassium hydroxide containing 20 g. of the alkali to 100 c.c. of water. Continue the boiling for three minutes. Pour off the alkali and wash the residue thoroughly with



Determining the percentage of wool in a fabric.

water. Put a drop of phenolphthalein indicator in the beaker and add dilute hydrochloric acid until the pink color disappears. This will neutralize the alkali left in the shreds. Wash thoroughly with water and dry as above at about 100° C. Cool and weigh again. Divide the loss of weight by the weight of the original piece of goods and you will have the percentage of wool.

The effect of strongly alkaline laundry preparations on woolen goods may be determined by immersing a piece of flannel in a hot solution of ordinary washing soda. Use a tablespoonful of the soda to a pint of water. Keep the solution nearly at the boiling temperature for half an hour. Then take the fabric out, rinse thoroughly, and examine to see what has happened to the goods. Washing soda is strongly alkaline and so are many laundry soaps and powders. It is used to soften water, but borax gives only a mildly alkaline reaction and for that reason is better.

To learn for yourself the difference between washing soda and borax in their effects on woolens, repeat the above test substituting the latter.

CHAPTER 15

MICHAEL FARADAY

SURELY you must have heard of Michael Faraday, the poor apprentice lad who sought to become assistant to the renowned Sir Humphry Davy. One day a generous friend gave Faraday a ticket to a course of lectures being given by Davy at the Royal Institution. The eloquence of Davy's delivery and the skill in demonstration, for which he was so justly renowned, aroused the enthusiasm of the young apprentice. Thenceforth he felt that business would forever be distasteful to him and that nothing would satisfy his heart's desire but the life of a scientist. Faraday had taken full notes of the lectures and made drawings of the apparatus used in demonstration. These he sent to the great Sir Humphry with a carefully written note telling of his ardent desire and requesting an appointment as laboratory assistant. Davy endeavored to discourage the young enthusiast from such a life but without avail, and then promised to give him a place at the first opportunity. Imagine Faraday's joy when a few weeks later a coach drew up at his humble lodgings and a liveried footman bore him a note announcing his appointment to the much coveted position.

And again we have an untutored young man, without previous experience, coming into a laboratory position of great responsibility and wonderful opportunity. Faraday's duties were to set up apparatus for Davy's lectures, assist him in performing the demonstrations, and still better to assist in making the discoveries which formed the basis of the lectures. Can you imagine a more fascinating beginning for the would-be scientist? Here he was, as if by magic, established in the best equipped laboratory of England and acting as assistant to the foremost scientist of Europe.

But Faraday improved his opportunities. He worked incessantly. He studied to obtain an education at night, and he experimented in every spare moment during the day. He went with Davy on a trip to the continent and there met the leading scientists of other lands. In a few years he was lecturing himself and developing that power of presentation which was to make him the most popular lecturer of his time. When Davy resigned as Director of the Royal Institution, Faraday became his successor and for more than forty years was known as one of the most distinguished scientists of any country. In those years he made some of the most notable discoveries of the last century. In electricity he laid the foundation for the marvelous developments of the present age.

And this life of wonderful achievement came from a boy's desire to experiment and become a scientist.

CHAPTER 16

STAINS AND BLEACHES

THE subject of stains and bleaches opens up for us a whole new realm of chemical investigation. Scarcely a week goes by that some boy does not come to my laboratory to have an ink stain or a grease spot removed. Very often a boy comes with gum on his trousers. One of his particular friends has contrived to have him sit on a wad of this apparently modern necessity. At such times a little knowledge of applied chemistry is very valuable.

Suppose we start with stains. A stain may be either physical or chemical. By physical I mean that no chemical change has taken place between the fabric and the substance causing the stain. Such stains are most easily removed. It is frequently impossible to remove a chemical stain.

For the removal of a large number of stains two solutions will be required. They are:

1. Tartaric or oxalic acids—50 g. to 250 c.c. of water.
2. Twelve and one half grams of ordinary chloride of lime to 250 c.c. of water. Boil this until a pink color appears, then filter and add 125 c.c. of

cold water. (Chloride of lime may be had at a grocery store, but be sure that it comes in a sealed metal can).

Keep these solutions in stoppered bottles. The first will keep indefinitely, but the second must be renewed at frequent intervals.

Ink stains are easily removed with these two solutions, especially if they are fresh. Try them both on fresh and old specimens of writing. Wet the ink with No. 1, take up the surplus liquid with a blotter, and then apply solution No. 2. Usually the ink will disappear. Ammonia will often remove fresh ink stains.

Try these solutions on *fruit, tea, or coffee stains*. After applying both solutions, rinse the material with water and, if necessary, apply them again. After the color has been removed, a yellow spot will often remain. In such a case, which occurs most often with woolen goods, apply hydrogen peroxide. This will oxidize the coloring matter and remove it.

Grease spots are probably the most common and troublesome stains. To remove them apply liberal quantities of gasoline, carbon tetrachloride, or benzene. Follow this treatment with soap suds and ammonia. If a grease spot is fresh, French chalk will usually absorb it. If not, use any one of the above solvents and then apply the chalk. (Carbon tetrachloride is "Pyrene," so much used in fire extinguishers. It is one of the best solvents known.)

Stains from dyes may usually be removed by appli-

cations of solutions No. 1 and No. 2. Chloroform is also an excellent solvent for these stains. So is carbon tetrachloride.

Gum may be removed by rubbing with carbon tetrachloride or chloroform. They loosen the sticky stuff and enable you to rub it off. Both evaporate very rapidly and leave no stain themselves.

Iodine stains are often very annoying to a laboratory worker, but they are easily removed. Ammonia water will usually dissolve them. A dilute solution of washing soda also does. If you have potassium iodide a ten per cent solution of it, that is, 10 g. to 100 c.c. of water, will do the trick. Denatured alcohol also dissolves iodine.

Spots of iron rust frequently occur on white goods. But, since iron rust consists of iron oxide which is soluble in dilute acids, the problem is easily solved. Your grandmothers covered the spot with lemon juice and salt and exposed it to the sun. The lemon juice contains citric acid, which dissolves the rust, while the salt serves to hold the liquid and keep the goods moist. A warm saturated solution of oxalic acid applied with a soft sponge will also remove these spots. *But remember that this acid is poison.*

Acid stains are the most annoying to a home-laboratory worker. About the only thing to do is to neutralize such stains with ammonia, and this should be applied just as soon after the acid falls upon the clothing as possible. If the acid has had time to do its work very little can be done. In the course of

many years of laboratory work I have ruined more suits of clothes than I like to remember. You will find that your best protection consists in wearing a laboratory coat similar to an automobile duster.

In case of paint either carbon tetrachloride or benzol will remove the spots. Turpentine followed with soap and ammonia is also good. Similar treatment may be used with *varnish*, but equal parts of wood alcohol, benzol, and acetone are better.

Mildew, and ink spots on rugs, may frequently be removed by an application of a solution of tartaric acid followed by "Javelle" water. The directions for preparing Javelle water will be found on the cans of chloride of lime which may be had from a grocery store.

BLEACHING

The bleaching of fabrics has become a very important branch of the textile industry. It is also often desirable to bleach goods in the household. But in order to remove a color you must produce a chemical change. The dye must be changed into a colorless substance. Furthermore this must be done without injury to the fabric. Sunlight is a great bleaching agent. You know that curtains and other colored goods often fade under the influence of the sun's rays. That is because the sunlight is able to produce chemical effects.

Bleaching Cotton.—The best bleaching agent for cotton is chlorine, but we seldom use chlorine gas

directly. The chlorine is absorbed in slaked lime and converted into bleaching powder, better known as ordinary chloride of lime. Obtain a can of this from the store and make a paste by stirring 2 g. of the powder into 100 c.c. of water. (Remember that unless bleaching powder comes in a sealed metal container, it is worthless).

Then prepare a very dilute solution of hydrochloric acid, about 1 c.c. of the concentrated acid to 25 c.c. of water. Also have at hand a solution of photographer's hypo containing 5 g. of the salt to 100 c.c. of water. Now cut some strips of colored calico and dip them into the paste of bleaching powder. After a moment or two place them in a small beaker containing the dilute hydrochloric acid. Continue this process, alternately dipping the calico into bleaching powder and acid until the color is removed. Finally wash the fabric thoroughly with water and dip it into the solution of hypo.

There is considerable chemistry involved in this process. The hydrochloric acid acts upon the bleaching powder to liberate chlorine. The chlorine then acts upon the water present and liberates oxygen. It is really the oxygen that does the bleaching. The hypo absorbs any excess of chlorine left upon the goods, for unless this is done the chlorine will weaken the fiber of the cloth.

Cotton goods may also be bleached directly by chlorine, but, since the preparation of this poisonous



Bleaching cotton fabric with chloride of lime.

gas in a small laboratory is somewhat dangerous, we had better not attempt it.

To bleach cotton on a larger scale stir thoroughly for some time the contents of a can of chloride of lime with 1 gallon of water and allow the mixture to settle. Pour off the clear liquid. Prepare equal volumes of the solutions of hydrochloric acid and hypo, using the same proportions as before. Also prepare a dilute solution of ordinary washing soda. Rinse the goods in the acid. Then pass them slowly through the solution of chloride of lime. Follow this treatment by dipping the goods first in the solution of washing soda and then in the hypo. Rinse thoroughly in water. If the color has not been entirely removed, put them through the process again. In some cases it may be necessary to leave the material in the chloride of lime for a short time. If a yellow color remains destroy it with hydrogen peroxide.

Bleaching Wool.—Since chlorine injures animal fibers, some other chemical agent must be employed. The substance that works best is our old friend sulfur dioxide. This gas, as you have doubtless discovered, has a strong odor, and it is also an important disinfectant.

Place two or three teaspoonfuls of sodium sulfite in a bottle provided with a stopper, to the under side of which you have pasted one end of a strip of colored flannel. Cover the sulfite with 30 c.c. of dilute hydrochloric acid (about 1 c.c. of the strongest acid to 10 c.c. of water). Then quickly insert the stopper so

that the strip of flannel will hang down in the bottle. Let the bottle stand for 24 hours. Then remove the cloth. You will find that the color has been bleached. The sulfite and acid liberate sulfur dioxide which in chemical language reduces the dye to a colorless compound.

Straw goods may also be bleached in this same way. In bleaching wool and straw, hydrogen peroxide may be substituted for sulfur dioxide.

Prepare a bottle of sulfur dioxide by the same method which you used in the chapter on Chemical Magic when you converted wine into water. Place in the gas a pink flower, a rose or carnation. In a few moments it will be bleached white.

To bleach indigo bubble sulfur dioxide from your generator through a little water for several minutes. Add to this solution a small quantity of zinc dust. Let the mixture stand for a time. Then add some of the liquid to a solution of indigo. It will be instantly bleached. If you have no indigo, substitute litmus solution.

A Dye that will not Bleach.—Not all dyes will bleach. Mineral dyes belong to this class. Prepare a solution of potassium chromate or dichromate. Immerse in the solution a strip of white cotton cloth. It is dyed yellow. But if placed in water, the color may be entirely washed out. It is not fast. Now prepare a solution of a lead salt. Either the nitrate or the acetate may be used. Dip the cloth first into this solution and then into the dye. Try to wash the color

out. You cannot. It is fast. The lead salt is what we call a mordant. It has united chemically with the dye to form an insoluble color. Mordants are very extensively used in dyeing, particularly in the dyeing of cotton goods.

Try the bleaching action of bleaching powder, sulfur dioxide, and hydrogen peroxide upon this color. This is the old and familiar household dye known as *chrome yellow* and is unaffected by ordinary bleaching reagents. Try separately strong ammonia water, sodium hydroxide, and hydrochloric acid on the color. What happens?

CHAPTER 17

JÖNS JAKOB BERZELIUS

BERZELIUS was a great Swedish chemist and for half a century was known as the "Czar of Chemistry." But he taught himself most of the chemistry he knew. When he went to college very little real chemistry was known. He had to experiment at home, and his room was a dark dingy place without even a window and only an open fireplace for heat. One day he obtained a gas from nitric acid which he suspected might be oxygen. He had heard of this gas but had never prepared it. His supply of chemicals and apparatus was very limited, and he had no one really competent to instruct him. But he arranged some bottles and collected this gas over water, just as we do nowadays. Then into one of the bottles he thrust a glowing splint and to his intense joy it blazed up brightly and lighted his dismal room with great brilliancy. He knew that he had prepared oxygen and his enthusiasm knew no bounds. Some day you will make an experiment that will thrill you with the joy of discovery.

Although Berzelius was to become the foremost chemist of Europe, his instructors in college regarded him as of very ordinary ability. Upon graduation the

professor in chemistry even told him that he deserved to fail. But Berzelius continued his chemical investigations. In his simple laboratories in Stockholm he prepared many new compounds and increased enormously the world's knowledge of chemical principles and processes. At length his fame spread abroad and chemists from other countries came to study under him. But they were often amazed to see the plain, simply furnished rooms and modest supply of apparatus and chemicals which sufficed for his great discoveries. Berzelius' work was done during the first half of the nineteenth century, and in the early part of this period college laboratories were unknown. It was only through the instruction given by such advanced students as this great Swedish chemist that chemistry progressed.

Nothing is more delightful than such a life of discovery and instruction. Even in your simple laboratory will be many conveniences and much apparatus which Berzelius did not have. And you, too, can make discoveries. They may not be new to the world, but they will be new to you. Some day you may make big discoveries, for there is still a host of them to be made.

CHAPTER 18

ELECTRIC FURNACES

IF you have a 110-volt source of current, either alternating or direct, no other piece of your laboratory equipment will afford you more pleasure and service than an electric furnace. And you can make one yourself. Electric furnaces are of two types—the arc and the resistance. One derives its heating power from an arc giving a temperature of about 3500° C. The other converts electric energy into heat through the agency of a metal wire of very high resistance.

We shall start with the arc furnace. The first arc light in history was produced by the renowned Sir Humphry Davy. Connecting a battery of 2000 cells to two pencils of carbon, he brought their ends together and then, withdrawing them gradually to a distance of 4 inches apart, he beheld a light of dazzling brilliancy and intense heat. In this arc he was able to melt or vaporize such substances as platinum, quartz, sapphire, magnesia, and lime. But Sir Humphry's arc light was only a curiosity. Dynamo current was unknown and electric cells were too expensive. The world had to wait for Faraday to discover the principle and for later scientists to apply it.

Making an Arc Furnace.—The materials needed for this furnace are fire clay, asbestos fiber, and water glass. These are not difficult to obtain. You will probably have to send to a chemical supply house for the fire clay, but it is not expensive. When these substances are mixed, they harden into a fireproof mass which is almost indestructible.

As a form in which to shape the furnace you will need a small box about 8 inches long and 4 inches square. I used a box in which blackboard crayons come. In the center of each end bore a hole big enough to take a standard electric light carbon. Then in an old basin mix fire clay, asbestos fiber (short shreds), and water glass solution until you have obtained a doughy mass. Pack a layer of this about an inch thick as firmly as possible in the bottom of the box. Next set a drinking glass in the center of the box and thrust through the end holes two wooden pins the size of the light carbons which you are to use later. Pack around these more of the mixture as tightly as possible until the box is completely full (Fig. 17). Smooth off the top and fill in the cavities with a mixture of fire clay and water glass alone. In the same way make a cover for the furnace of the same size and about an inch thick. Set the box and cover in some warm place to dry for about a week. The top of the furnace is a good place. When it is dry and hard, break away the box, pull out the pins, and retouch the whole with a paste of fire clay and water glass. Strips of asbestos paper glued to the outside improve the ap-

pearance. Insert light carbons to which you have attached copper lead wires, and your furnace is ready for action.

Making a Rheostat.—But such a furnace as this will blow the electric fuses of your lighting circuit, unless a suitable resistance is connected in series with it. A rheostat you will find necessary for much of your

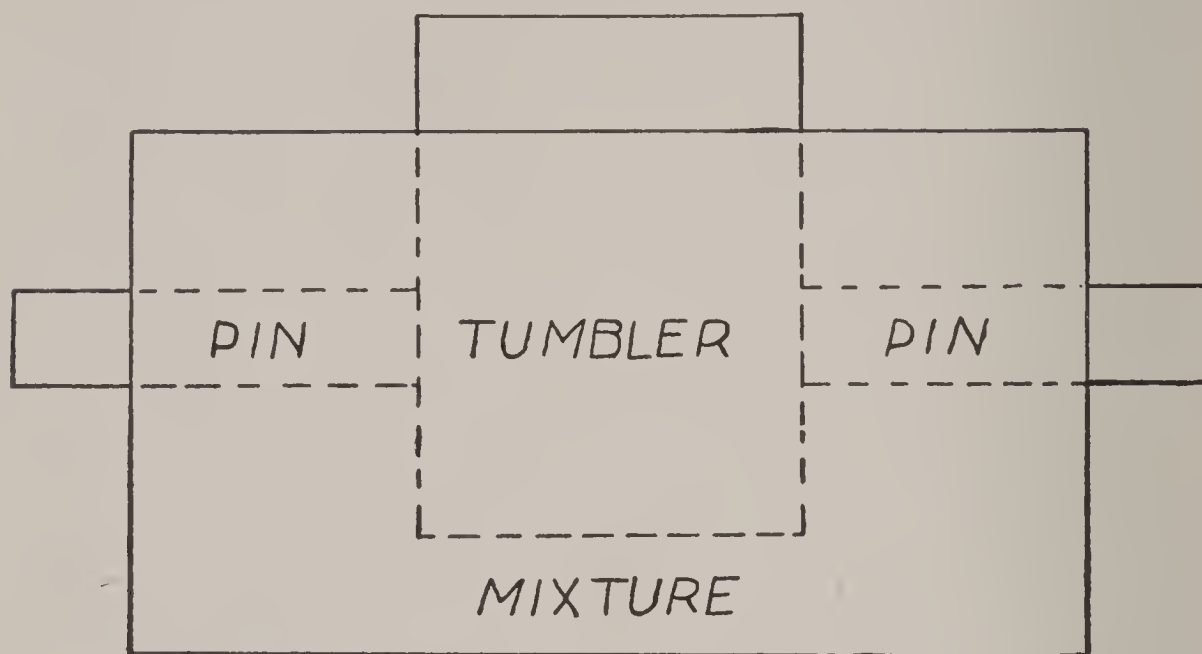


FIGURE 17.

The arc type of electric furnace.

electrical work, so we might as well start about making one first as last. A lamp-board rheostat is the cheapest, and the easiest to make as well as one of the most serviceable. A dozen lamps will give you a wide range of current control, and should you need more current, additional lamps can be added. The accompanying figure shows the grouping for a board of 18 lamps. If you can find old 16 candle power carbon lamps,

they will give you a larger amount of current (Fig. 18).

To use this board connect your apparatus, whatever it may be, in series with the lamps, as indicated. Unless you know about what current you will need, have

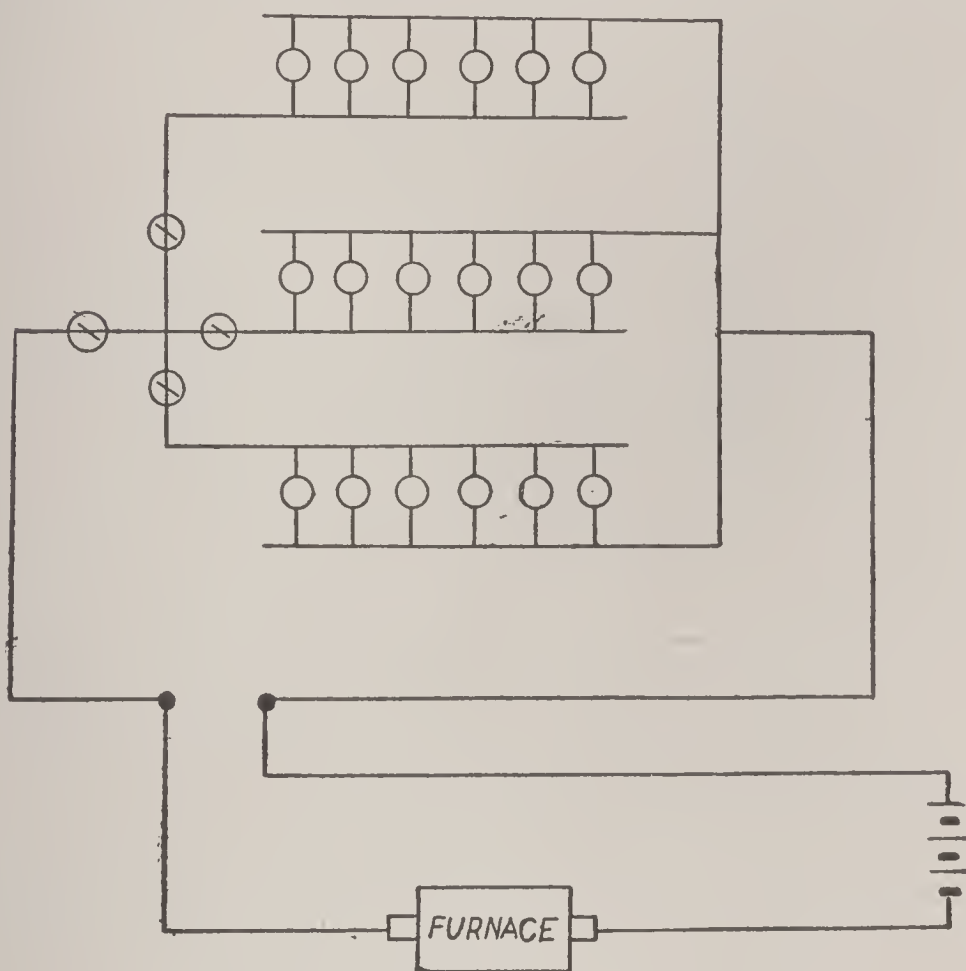


FIGURE 18.

Diagram of lampboard rheostat showing electric furnace connected in series with it. After the arc has been "struck," it is safe to short circuit the furnace and cut out the lamps, if the lighting circuit fuse is as high as 15 amperes. To do this put a heavy wire across the main binding posts.

all the lamps screwed loose from their sockets at the start. Then turn on the main switch and one of the branch switches. Screw in one lamp. That will give you a small current, something less than half an ampere. It will never be enough to injure any ap-

paratus that you may have. Screw in another of the lamps on that circuit and you will nearly double the current. Continue to do this until all of the lamps on that branch have been turned on, or until you have obtained sufficient current for your purpose. If one branch does not give enough current, turn on the second adding one lamp at a time, as before. You will seldom need to use all of the 18 lamps.

Of course you can buy a rheostat from a supply house, but to get a good one will probably cost you as much or more than the one just described. By winding resistance wire, either Nichrome or German silver, on a wooden core well wrapped with asbestos, you may make a rheostat that will be very serviceable. One of Nichrome wire is the best rheostat for an arc furnace. A rheostat of fairly good capacity may be made from 210 feet of No. 26 German silver wire or 80 feet of No. 20 Nichrome wire. Use the wire bare and wind the turns as closely together as possible. By mounting this on a baseboard between two end-pieces and providing a sliding contact after the fashion of a radio tuning coil, you will be able to obtain a variable rheostat having a wide range of resistance. Attach one end of the wire to a binding post, and make the other connection through the sliding contact (Fig. 19).

Using the Furnace.—First you must learn to “strike an arc.” Connect the furnace in series with your rheostat and, if it is of the lamp-board type, turn on nearly all of the lamps. Then bring the ends of the carbons

together and quickly withdraw them about a quarter of an inch. An arc of dazzling brilliancy will be produced, so bright that for the protection of your eyes *you should wear colored glasses*. If it is too bright, turn off some of the lamps. To obtain more heat turn on more. After the arc is struck you may cut out the rheostat entirely.

By means of tongs hold in the arc such metals as iron, copper, and aluminum. Note how readily they

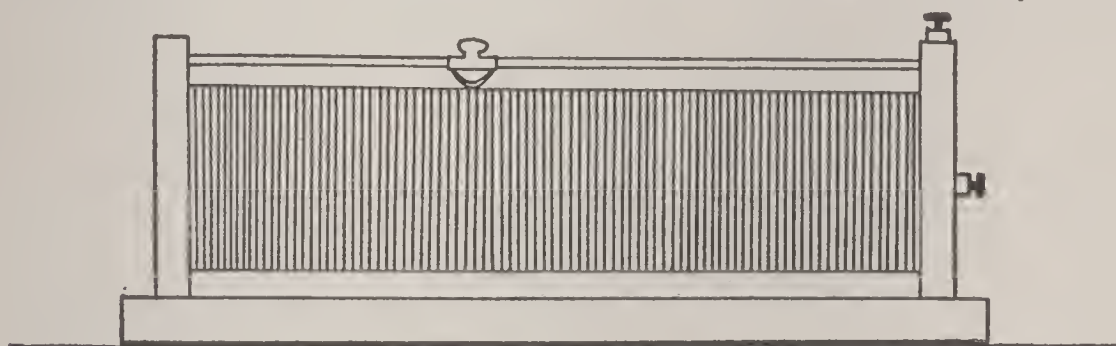


FIGURE 19.

A rheostat made by winding German Silver or Nichrome wire on a wooden core wrapped with asbestos. One end of the resistance wires is attached to the center binding post. The number of turns used is regulated by the sliding contact.

melt. A lump of quicklime may be heated to brilliant incandescence. Place just beneath the arc a small fire-clay crucible containing some pieces of sheet copper or copper wire. In a few minutes they will melt, and to melt copper requires a temperature of nearly 1100°C .

A Resistance Furnace.—Interesting as an arc furnace is, it is not nearly as useful as a resistance furnace. This type of furnace depends upon the electrical resistance and heat producing properties of an alloy known as Nichrome wire. For a long time the use of this

wire was covered by patents, and amateurs could not purchase it, but it is now available.

We shall make first a crucible furnace. For this you will need, as before, fire clay, asbestos fiber, and waterglass. In addition you must have a corrugated Alundum core, $2\frac{1}{2}$ inches deep and 2 inches in diameter, some Alundum cement, and 42 feet of No. 20

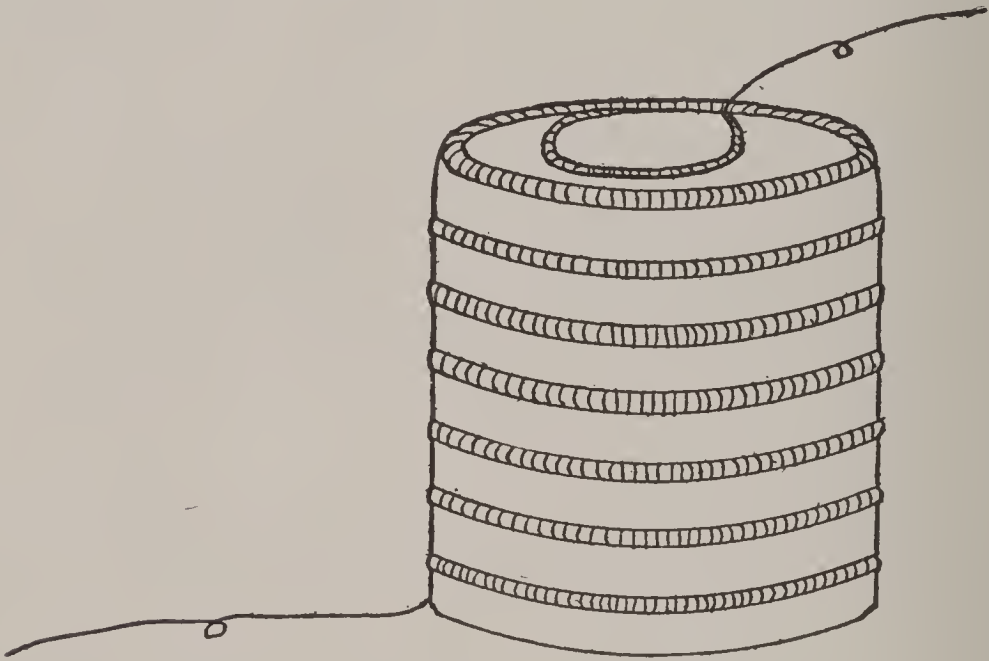


FIGURE 20.

Alundum core for crucible type of electric furnace showing Nichrome wire in place. Cement must now be added.

Nichrome wire. The first step is to wind this wire as closely as possible on a spindle $\frac{3}{32}$ of an inch in diameter, leaving 1 foot free at each end for connections. In making my furnace I got the wire wound in a machine shop. Stretch the coil slightly so that the turns will not touch and then wind it about the Alundum core, fitting it into the grooves (Fig. 20). By means of cords or rubber bands hold the turns in place,



Lampboard rheostat connected in series with an arc furnace. A resistance furnace is shown at the left.

until you can cover the whole with Alundum cement to a depth of about a quarter of an inch. Put it in a warm place and let the cement set.

Now obtain a box similar to the one used in making the arc furnace. Into the bottom press a layer of fire clay, asbestos fiber, and waterglass. In the center of this set the Alundum core and pack firmly about it more of the mixture until the box is level full. Bring the wires out to one end and secure them to binding posts set into the plastic mass. Cover the top with a layer of Alundum cement and make a cover, as you did for the arc furnace, but this time leave a small hole for a vent over the center of the pit. Then connect the lead wires to a 110-volt source of current and let the heat of the furnace itself dry it out. But remember that you cannot obtain a high temperature until all the moisture has been driven out. No additional resistance will be needed with this furnace, and you will find that, when it is dry, the crucible may be brought to a white heat. But first remove the box and smooth the outside with cement.

Using the Crucible Furnace.—In using this furnace always set it upon a piece of asbestos board. To get up the maximum heat will require from 20 to 30 minutes. The materials to be heated are put into a small crucible which is set in the fire-pit. In order to obtain the greatest heating power it will be necessary to keep the fire-pit covered.

Making Brass.—An excellent test for your furnace will be the stunt of making brass, for to melt copper

requires a temperature of 1084° C. If you can do this, your furnace is well made. Weigh out into a small fire clay crucible 35 g. of copper, place the crucible in the pit, and bring the furnace to its highest temperature. At first no doubt the copper will seem very stubborn. Although white hot, it will not melt. Yet these metals have the peculiar property of remaining in solid form, until a certain temperature is reached. Then of a sudden, when that temperature is reached, the metal quietly turns to liquid as simply as a piece of butter in a heated spoon. Presently, as you remove the lid, you will behold a beautiful mirror of molten copper.

When the copper has melted, dissolve in it 15 g. of zinc. After allowing a few moments for the two metals to mix, seize the crucible with a pair of tongs and plunge it into a pail of water. There will be a vigorous seething of the water, a dense cloud of steam, and then the metals will have frozen.

You have formed an alloy. An alloy is a mixture of metals melted together and possessing properties often differing widely from those of the constituent metals. Alloys are tremendously important. Bronze, an alloy similar to brass, was known long before iron and steel. The early forests of Europe were felled with bronze axes, and bronze implements of warfare were used to attack the enemy and repel invasion. Nowadays we can make alloys to meet any requirement of science or industry. The wire in your furnace is an alloy. So are solder, shot, German silver, and

coin metals. You may prepare other alloys in your furnace but many of them do not require so high a temperature.

Making Quicklime.—Fill your crucible with lumps of marble or limestone. (To test for limestone, put a drop of acid upon it. If it effervesces, i.e., gives off a gas, it is limestone. If not, it is probably sandstone.) Let the marble heat at the highest temperature of the furnace for from one to two hours. Then let it cool. You have driven off carbon dioxide from the marble and quicklime is left. This is the same substance that is prepared in limekilns on an enormous scale. Quicklime is one of the three most important raw materials of chemical manufacture.

When the product of your furnace is cold, place a few of the lumps in a small basin and let fall upon them a few cubic centimeters of water. If the lime is well burned, it will immediately begin to get hot, steam will form, and the hard mass will swell and crumble into a powder. This process is called *slaking*, and it is what you frequently see taking place in a mortar box on the street. The water put on the lime is cold, and the heat is developed by the chemical action.

A Beehive Coke Oven.—You can turn your furnace into a beehive coke oven. Fill the crucible with quite small pieces of *soft* coal. Bring the furnace to full heat, set the crucible in place, and put on the cover. In a few moments you will be able to light the gas which escapes through the vent in the cover and it will

burn with a yellow smoky flame. In an hour you may allow the furnace to cool and remove the product. You will find that the soft coal has swollen into a light steel-gray substance. It is almost pure carbon, and coke is another of the three most important raw materials of chemical manufacture.

The Metallurgy of Copper.—Fill the crucible with a mixture of powdered charcoal and the wire form of copper oxide. You can prepare this copper oxide by heating a crucible full of short lengths of copper wire for an hour with occasional stirring. Put the mixture of oxide and charcoal into the furnace and heat it for a half hour. Then remove the crucible with tongs and pour its contents into a large beaker or basin of water. Hold the beaker under the faucet and let water run into it until all of the charcoal has been washed out. The copper is heavy and will remain in the beaker. If you do not have running water, throw water into the beaker from a pail. The charcoal has reduced the oxide to bright metal.

Making Charcoal.—This time fill the crucible with small pieces of soft wood. Match sticks will serve. Again you will be able to light the gas that issues from the crucible, and the product will be charcoal.

Another Type of Resistance Furnace.—There is another electric furnace called the combustion furnace, which some of you may be able to make. The materials for this are the same as for the crucible furnace, except that you will use a long hollow Alundum core open at both ends. This should be about 12 inches

long and grooved like the other. You will need a somewhat longer length of Nichrome wire, enough to wind the core after it has been made into a coil as before. Seventy-five feet will be sufficient. After putting the coiled wire in place, cover it with Alundum cement and let it set. Be sure the ends of the wire are exposed for connections. Then set the core inside a long cylindrical box and pack about it a mixture of the fire clay, asbestos fiber, and waterglass. Dry it with the heat of the electric current, and, after removing the core from the box, smooth off the outside with Alundum cement. Mount it on legs made of strap iron.

The heat of this furnace will be too great for ordinary glass combustion tubing. A silica combustion tube will be required, and this you must obtain from a chemical supply house. It should be about 18 inches long. This furnace, too, will require no rheostat in series with it. It will take the place of the 4-tube burner and may be used for a number of important experiments.

Preparation of Water Gas.—Throughout the middle portion of the silica combustion tube put some pieces of charcoal. With rubber stopper, glass and rubber tubing connect to one end of the combustion tube a flask containing water and mounted on the ring-stand over the Bunsen burner. To the other end connect a delivery tube leading to inverted bottles of water in your pneumatic trough. The set-up aside from the furnace will be identical with that used in the prepara-

tion of hydrogen by passing steam over hot iron filings. As the heat of the furnace comes up, bring the water in the flask to a boil and collect the gas which escapes from the delivery tube in the bottles. The bottles will at first be filled with a cloud of vapor, but this will soon disappear, and you will have a colorless gas. The hot carbon has taken the oxygen from the steam forming carbon monoxide and leaving the hydrogen of the water in the free state. The gas consists, then, of a mixture of carbon monoxide and hydrogen, both of which are combustible.

To test the water gas invert one of the bottles upon the table and immediately apply a match. The gas will burn with the characteristic blue flame of carbon monoxide. Pour into the bottle a little limewater and shake it about. You will obtain a white precipitate showing that carbon dioxide has been formed in the combustion. A film of condensed vapor on the sides of the bottle will prove the presence of hydrogen in the gas.

The preparation of hydrogen by passing steam over hot iron filings, already described in a previous chapter, may be carried out beautifully with this electric furnace.

The preparation of copper oxide from short lengths of copper wire may be much more effectively done in this furnace than in the crucible type. Place the wire in the combustion tube and after bringing the furnace to a good heat pass over the wire a stream of oxygen from your generator.

Other uses of the furnace will suggest themselves. Quicklime may be made in it. But it is not best to prepare coke in it, for the soft coal swells in the process and it would be difficult to remove it from the tube.

CHAPTER 19

SOME METALS AND THEIR ALLOYS

I AM not going to forget the fellow who cannot make an electric furnace. Many interesting experiments with metals and their alloys may be done with a Bunsen burner or alcohol lamp.

You do not need to be told that metals from earliest times have been among the most useful of the elements. Just as men have increased their knowledge of them, so has their mastery of Nature grown. Gold was one of the first metals known, doubtless because it occurs in the free state. Its metallurgy, the process of extracting it from the ore, is easy. And, if you have read history, you know that there was a long period of many centuries known as the *bronze age*. Primitive men first used implements of the chase and warfare fashioned from stone and flint. Then some cave man, raking over the embers of his dying fire, probably discovered shining globules of metal. Without doubt they were copper or tin, for these metals are most easily reduced from their ores. At first this incident only aroused his curiosity. The bright metal appealed to his fancy, and he probably used it only for ornaments. Not for many years, possibly centuries, did it occur to him that this new material could be of any practical value. And quite likely a long period

elapsed before he made any deliberate attempt to duplicate the process, which he had accidentally discovered. Slowly he learned that these metals possessed properties very much superior to those of stone and wood. From them he began to fashion crude weapons. How he blundered onto the fact that copper and tin together form an alloy surpassing in usefulness either metal alone, we do not know. Painfully he forged ahead. At some time in the gray mists of antiquity, he discovered the metallurgy of iron, and then he began to dominate the earth. Today steel is the symbol of power. Because metals touch our lives at a thousand points, we should know something about them.

What Are Metals? Chemically we distinguish between metals and non-metals. *A metal* is an element whose oxide with water forms a base. *A non-metal* is an element whose oxide with water forms an acid. In addition metals have certain characteristic physical properties. In the pure state they have a high luster and are good conductors of heat and electricity. Many of them are tough and malleable. Some are brittle. Chemically metals vary from the intense activity of potassium and sodium to platinum and gold, which are little affected by the strongest reagents. Some will burn, others will not. Some melt at low temperatures, others only in the heat of the electric arc.

Chemical Test for a Metal.—Cut a bit of sodium or potassium. Note its bright metallic luster and the rapidity with which it tarnishes. These are the most

active metals known. So strong is their affinity for oxygen that they unite with it immediately upon exposure to the air and for that reason must be kept under kerosene. Throw the bit of metal upon some water in an evaporating dish. Nothing could illustrate better its very active nature than the ease with which it decomposes this very stable compound. Dip into the solution a piece of red litmus paper and note the deep blue color, showing the presence of a base. This is the test for a metal.

As you know, limewater, which is slaked lime in solution, also turns red litmus blue. It contains in chemical combination the metal calcium. Sometimes the base of a metal such as that of iron is insoluble and will not change the color of litmus, but it will neutralize an acid to form a salt, and that test is equally good.

Heating Metals in the Air.—In no other respect do metals exhibit more marked differences than in their behavior toward heat. Some metals burn, some oxidize, while still others are wholly unaffected. The very usefulness of metals depends upon these varying properties.

Hold a piece of magnesium ribbon in the flame and note its brilliant combustion. At that temperature its affinity for oxygen is exceedingly strong. The product is a white powder that you can crush in your fingers. It in no way resembles the original substance. Then, if you have it, hold a platinum wire in the flame. It glows brightly, but that is all. When it has been re-

moved and cooled, its properties are exactly what they were before. No chemical change has occurred. We might repeat the experiment and thus discover that the properties of metals differ widely under the influence of heat.

Does a Metal Grow Heavier, when Heated in the Air? Try it. Place in a porcelain crucible one or two grams of granular tin and exactly counterpoise it in your hornpan balance with lead shot. Then, placing a clay triangle on the ring-stand support, mount the crucible over the Bunsen flame. As the tin melts, a dark scum will form on the surface. With an iron wire pull this to one side, thus exposing a fresh surface to the air. Again this scum will appear, and the process may be repeated over and over. Gradually you may change all of the tin into this gray powder. When you have done so, allow the crucible to cool and then replace it in the balance. If you have been careful not to lose any of it, the powder will be considerably heavier than the original tin. It is no longer tin but tin oxide. And yet to understand the nature of this process required long years of patient experimentation. What is true of tin is true of all metals which will unite with oxygen.

Making Solder.—Solder and pewter are two of the easiest alloys to make because lead and tin, from which they are prepared, have low melting points. Into a small but rather deep fire clay crucible weigh equal quantities of these two metals. Mount the crucible over the Bunsen burner and starting with a small flame

gradually increase the temperature until all of the metal has melted down into a uniform mixture.

For shaping these alloys into sticks, a plaster of Paris mold will be convenient. Make a mixture of plaster of Paris and water and fill with it a long narrow box cover. While it is still soft, press into it two or three lead pencils. The molten solder may be poured into these depressions. When the sticks are hard, you may put them into small labeled specimen tubes and add them to your chemical museum.

An Alloy that Melts in Hot Water.—An alloy that will melt in hot water used to be only a curiosity, but now it is of great importance for filling the perforations in the pipes of automatic fire extinguishers of the sprinkling type. This alloy is known as Wood's Metal and is made by melting together 80 g. of bismuth, 40 g. of lead, 20 g. of tin, and 15 g. of cadmium.

After the sticks of the alloy have hardened, heat some water in a beaker to about 70° C., and, holding one of the sticks with the tongs, thrust it into the water. It will quickly melt and run about the bottom of the beaker like mercury. Pour cold water into the beaker and the metal will immediately freeze.

Such low melting point alloys are also used in making fuse wire for the protection of electric circuits.

Lead is hardened for use in shot by alloying it with arsenic. Melt 100 g. of lead and while it is molten sift in a half gram of arsenic—not white arsenic but the element arsenic. Compare the properties of the alloy with that of soft lead.



Melting an alloy in hot water heated over an alcohol burner.

Type metal, an alloy which has the peculiar property of expanding upon solidification, may be prepared by melting together 75 g. of lead, 20 g. of antimony, and 5 g. of tin.

You may galvanize iron by dipping it into molten zinc. Clean a large "cut" nail by heating it to as high a temperature as possible in your Bunsen burner, allowing it to cool, and dipping it in hydrochloric acid. Then scour it with sand, wash it off, and reheat. While it is still hot, dip it into olive or cotton-seed oil and then into a crucible containing molten zinc. Upon removing and cooling you will find that the iron has been galvanized.

By cleaning sheet iron and dipping it into molten tin, *tin plate* may be prepared in a similar way.

Home-made thermit will afford you one of the most interesting and spectacular demonstrations that can be performed with metals. As you may know, *thermit* is the trade-name of a mixture which gives one of the highest temperatures known to science. Its great use is in welding iron and steel. Its ignition gives liquid iron at a temperature of more than 3000°C .

Mix about equal quantities of aluminum dust and ferric oxide. Place a considerable quantity of the mixture in a small dish made of thin sheet iron, or in a small tin basin. Prepare some ignition powder by mixing equal parts of magnesium dust and powdered potassium chlorate. In a little depression hollowed out of the top of the heap of thermit place a half tea-

spoonful of the ignition powder and insert in it a short length of magnesium ribbon (Fig. 21).

Place the dish and mixture on a tripod support setting in a box containing an inch layer of sand. Light the magnesium ribbon and very quickly the reaction will spread throughout the mass of powder, producing a light of dazzling brilliancy and liberating white-hot

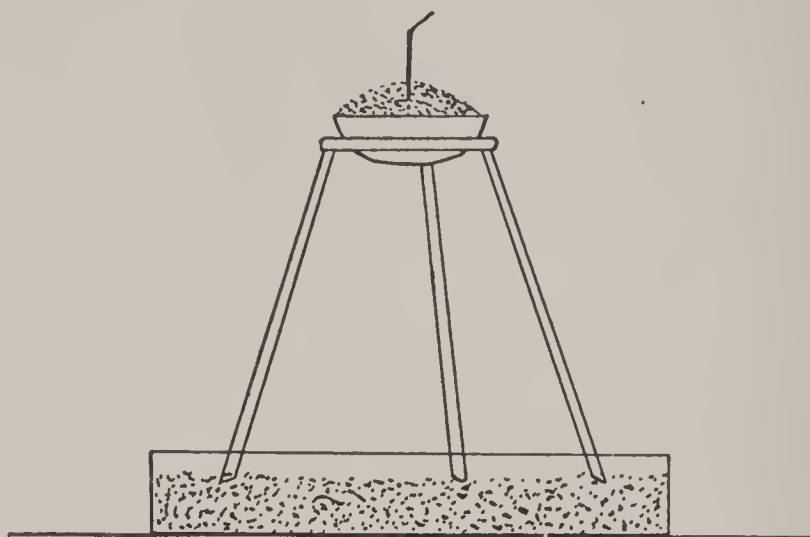


FIGURE 21.

A mixture of thermit ready for ignition.

molten iron, which will melt a hole through the iron dish and flow like lava into the sand box below.

In this experiment it is well to protect your eyes with colored glasses.

If you can find a 2-inch length of hollow tiling about an inch and a half in diameter, rest it upon a thick sheet of soft iron or steel and fill it with a fresh mixture of thermit, with the ignition powder and fuse at the top. Upon ignition you will obtain a well of molten metal which will fuse with the iron beneath

and weld itself to it. When the mass has cooled, break away the tile and you will find that the "boss" of iron cannot be removed by the most vigorous hammering.

Thermit is very widely employed for welding broken parts of locomotives, street car rails, propeller shafts, and any of the other multitudinous breaks in iron and steel. If you do not wish to prepare it, real thermit and ignition powder may be purchased from a chemical supply house.

The "Alchemy" of Metals.—To the alchemist the transformation of one metal into another seemed a simple thing. He thrust a bar of iron into a solution of blue vitriol and behold presently he had copper. To him, ignorant as he was of real chemical processes, it seemed perfectly plain that the iron had been changed into copper. But we know now that some metals will replace others from their solutions. That is because some metals are more active than others. In fact all the metals may be arranged in an order such that any one of them may be replaced from its solution by any metal preceding it and it will replace any metal following it.

In two sets of cylinders or tall test tubes place fairly strong solutions of the nitrates of zinc, lead, copper, mercury (either mercurous or mercuric), and silver. Cut some long narrow strips of sheet zinc and thin copper foil. Into one of the two sets of test tubes containing these solutions thrust strips of zinc and into the other strips of copper. Almost immediately

the action will begin, but you will not obtain deposits in every case. Note carefully those in which you do. Of course zinc will not replace itself in zinc nitrate, but in the case of all the other metals you will obtain a deposit, showing that of this group zinc is the most active. In the course of a half hour you will observe that the copper solution (copper sulfate may be substituted for the nitrate) has been decolorized and that a heavy deposit of metallic copper has formed in the test tube. The copper may be dark from impurities in the zinc, but you may melt it in your electric furnace into a metallic button. Of course at the same time that the copper is replaced the zinc goes into solution.

Examine the strip of zinc placed in the solution of mercury. Rub it and note the silver luster. Break the strip and examine the broken edge. The zinc is now brittle and the edge is white clear through, showing that the mercury from the solution has penetrated the metal.

Remove the spongy deposits of silver from the two test tubes containing solutions of this metal and rub them on a glass plate. The bright luster of the white metal will be obtained. You may redissolve the metal in nitric acid and obtain silver nitrate again.

Pure Silver from a Silver Coin.—Dissolve a small silver piece in not too dilute nitric acid. You will obtain a solution with a blue color, due to the presence of copper, for to harden coin metal an alloy of about 92 per cent silver and 8 per cent copper is employed. Your solution will contain both silver nitrate and

copper nitrate. The problem is to separate pure metallic silver from the solution. The preceding work on the replacement of metals will enable you to do this. No doubt some of you see the method already. All you need to do is to insert a metal that will displace the silver but not the copper. Zinc cannot be used, for it will displace the copper as well as the silver. But how about copper itself? You have just seen that copper will drive out silver from its solution, but of course it cannot displace itself.

Thrust a strip of copper into the solution and allow it to remain until you are sure all of the silver has been displaced. Remove the copper and scrape all the silver off onto a filter paper placed in your funnel. Into a clean beaker pour off the clear copper nitrate solution, leaving in the bottom any sediment that may be left. This sediment will of course be silver. Throw it upon the filter paper with the rest of the silver, using distilled water to remove the final portions. Now wash the silver by passing distilled water through the filter for some time. Use only small amounts of water at a time, and, if you wish perfectly pure silver, continue to wash it until the wash water that comes through the filter will give no blue color when a few drops of ammonia water are added to it. You may melt this silver down into a small button in your electric furnace, or redissolve it in nitric acid and obtain a solution of pure silver nitrate. If you wish to obtain crystals of silver nitrate, evaporate off part of the water and allow the rest of the solution to cool.

By a similar process of crystallization you may separate the copper nitrate.

The principles involved in this replacement of metals are at the bottom of a great deal of chemical work. The chemical action of electric cells and the electrolysis of metals are examples.

CHAPTER 20

JUSTUS VON LIEBIG

COMING upon the scene a little later than Berzelius and extending his period of service just a quarter of a century longer was Justus von Liebig, a great German chemist. He was an old autocrat and made many enemies as well as warm friends. As a lad he was apprenticed to an apothecary, and apothecary shops in those days were small research laboratories, that is, laboratories in which original experiments and discoveries were made. Liebig's bent, like Davy's, seemed to be for making explosives. One day young Liebig saw some traveling showmen prepare fulminate of silver, a very explosive compound. This of course appealed tremendously to the fancy of the young experimenter, and he immediately attempted to duplicate the process. The result was a terrific explosion which tore a hole through the roof of the apothecary shop and blew Liebig out of the drug business, for his employer dismissed him forthwith.

But Liebig had enjoyed too much of the pleasures of chemistry to permit his fondness for experimenting to be thus rudely destroyed. Therefore he asked his father to send him to a university where he might study chemistry to his heart's content. His father consented

and off he went, first to a German university and then to Paris. At Paris he studied in the laboratory of Gay-Lussac, a famous French chemist, just as others had studied with Berzelius.

Returning to Germany, Liebig established the world's first college laboratory for actual instruction in chemical work. Students flocked to him from every part of Europe and America. And he was a wonderful teacher. Some of the most distinguished chemists of the last century were his pupils.

Liebig invented new processes of organic analysis. He founded the science of agricultural chemistry, and among his discoveries of new compounds is that of chloroform. And possibly it should be mentioned that "Liebig's Beef Extract" is a product of his creation. It was the happy fortune of Liebig to live and work in the days when chemistry was in its youthful prime, and he made the most of his opportunity.

CHAPTER 21

MORE ABOUT THE CHEMISTRY OF COMBUSTION

PROBABLY your definition of combustion would be the chemical union of oxygen with a substance to produce heat and light. But we may have combustion without oxygen. Any chemical change accompanied by light is a form of combustion. If there is light there will always be heat. In fact there can be no chemical change without the giving off or the absorption of heat. But can there be combustion without a flame? Your first impulse will be to say "No." A little reflection will show you that there may be. The term *flame* is the name we give to the combustion of two gases or of a gas and a vapor. A vapor is the gaseous product obtained from a substance which at ordinary temperatures is a liquid or a solid.

Combustion without a flame.—Place a little sulfur in the bottom of a test tube and bring it to a boil in the flame. When the hot sulfur vapor is issuing from the tube drop into it a strip of thin sheet copper or some fine copper wire. You will note that the copper glows brightly. It burns, but there is no flame. When the test tube has cooled, remove the strip. It is no longer metallic copper. Its bright luster is gone. It

is no longer flexible. You can break it easily in your fingers. You now have a bluish-black substance called copper sulfide. In the union of the copper there were heat and light, but no flame.

Mix one part of iron filings with twice that bulk of finely powdered sulfur. Place this mixture in a test tube and heat in the flame until the contents of the tube begins to glow. Withdraw the test tube and note that the combustion which has started spreads throughout the whole mass. Again you have combustion, but no sign of a flame.

Mix thoroughly equal parts by volume of finely powdered sulfur and zinc dust. Place this mixture upon a square of asbestos board and ignite it by throwing the Bunsen burner flame down upon the heap. You will obtain a vivid combustion and, it may seem, a flame, but there is none.

Prepare a bottle of oxygen and thrust into it with your deflagrating spoon a glowing piece of charcoal. The glowing will be much more brilliant, but you cannot detect a flame. Repeat the burning of steel wool in oxygen and try to observe a flame. Even in the very vivid combustion of magnesium ribbon or flash powder, there is no true flame.

Mix thoroughly equal parts of copper oxide in powdered form and finely powdered charcoal. Place them in a *hard glass* test tube fitted with a one-hole stopper and delivery tube. Shake the tube so that the mixture becomes spread out somewhat in a thin layer and clamp it to your ring-stand in a nearly horizontal position.

Let the delivery tube dip into a test tube containing a little limewater. Then heat the mixture. Carbon dioxide will be driven off which may be detected by the white precipitate obtained in the limewater. After a short time, as you strongly heat the tube, the mixture will begin to burn and a bright glow will spread throughout the mass. Combustion is taking place. The copper oxide is giving up its oxygen to the charcoal so rapidly that it is being heated to incandescence. But you see no flame.

There are other illustrations of combustion without a flame but these are sufficient to demonstrate the point.

The Bunsen Burner Flame.—Light your Bunsen burner and open the air holes at the bottom of the tube. Note that the flame consists of two cones—an inner and an outer. There is really a third consisting of an invisible envelope of burned gases surrounding the other two. The *inner* cone consists of unburned gas, cool at the bottom but reaching a temperature at the tip way above its kindling point. To show that the gas in this cone is not burning, quickly thrust into it at the bottom a match head. The match will not ignite, and you may carefully raise it nearly to the tip of the cone before it will do so. Holding a glass tube in a nearly vertical position thrust it into this inner cone and light the gas as it escapes from the upper end. The tip of this inner cone is called the *reducing flame* because the hot unburned gas in it has a great affinity for oxygen and, if thrown upon a me-

tallic oxide, will reduce it. The gas in the *outer cone* is vigorously combining with oxygen and this portion is known as the *oxidizing flame*. As we shall see, both of these flames have important uses in chemical analysis.

Burning Air.—We usually think of air as the supporter of combustion and gas as the combustible substance. You will recall, too, that hydrogen will burn, although a lighted candle thrust up into an inverted bottle of the gas will go out. Still these terms—supporter of combustion and combustible substance—may be inverted. We may burn air in an atmosphere of illuminating gas.

Fit a lamp chimney with a 2-hole stopper carrying a glass elbow and a straight glass tube about 1 centimeter in diameter and reaching well up into the chimney. Connect the elbow with the gas supply and over the top of the chimney place a piece of asbestos having a hole in the center. Placing your hand on the top of the asbestos, turn on the gas and light it at the lower end of the straight glass tube. Immediately remove your hand and light the gas at the top of the chimney too. As you do so, you will observe that the flame passes from the bottom of the tube to the top and continues to burn inside the chimney (Fig. 22). Air drawn in at the bottom of the tube is burning in an atmosphere of illuminating gas. In this case the air is the combustible substance and the illuminating gas the supporter of combustion.

Protecting Gas from Ignition.—With the tongs hold a 4-inch square of rather fine mesh iron or copper gauze about 3 inches above the Bunsen burner. Turn on the gas and light it above the gauze. You will find that, while the gas will burn above the gauze, the flame will not strike down through and ignite the gas below. This is because the metal is a good conductor of heat.

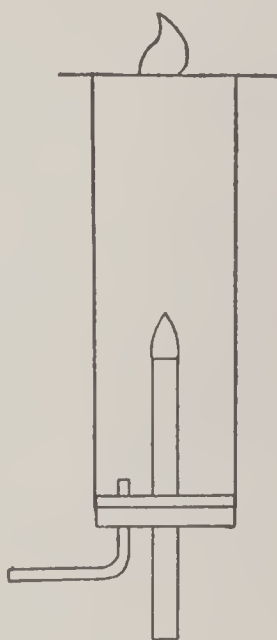


FIGURE 22.

Apparatus for burning air in an atmosphere of illuminating gas.

It distributes the heat over so large a surface that the gas cannot be brought to its kindling temperature.

This is the principle employed so successfully a century ago by Sir Humphry Davy in his invention of the miner's safety lamp. He surrounded the naked flame with a cylinder of copper gauze. When the miner enters a region containing the dangerous mixture known as "fire damp," the explosive gases penetrate the gauze causing a series of small explosions and thus

give a warning which leads him to beat a retreat. The gauze does not permit the flame to pass through and ignite the gas outside. Thus explosions, which would wreck the mines, are frequently prevented.

You Can Easily Make a Safety Lamp.—Roll a sheet of fine iron or copper gauze into the form of a cylinder and fasten it along the edges. By means of wire supports secure a candle in one end and light it. Now holding the cylinder in a nearly horizontal position over the Bunsen burner, turn on the gas, but the burner will not light. The gas will pass through and burn on the inside of the cylinder, but there will be no flame on the outside.

In order that a gas may undergo combustion a certain definite temperature, called the *kindling temperature*, must be reached.

Another very striking experiment to illustrate this principle may be shown with a funnel, a square of gauze, and some gunpowder. Clamp the funnel stem down to the ring-stand support and connect it to the gas supply. Place over the mouth of the funnel a fine mesh gauze and place in the center of it a little heap of gunpowder, being careful not to let it spread out very far. Turn on the gas and light it. You will have the quite startling spectacle of gunpowder being immersed in the heart of a flame and yet unable to burn (Fig. 23).

An Acetylene Lamp.—You may make a lamp for studying the combustion of acetylene gas and also obtain a beautiful and useful light by putting together



Burning air in an atmosphere of illuminating gas. Gas is burning at the top of the chimney and a small jet of air inside the chimney.

the following materials—a wide-mouth bottle of about 2 quarts capacity, a good-sized lamp chimney with straight sides and constriction toward the bottom, small disk of coarse gauze, one-hole rubber stopper, 6-inch length of glass tubing, rubber connection and pinch cock, acetylene gas tip, and 18 inches of heavy aluminum wire.

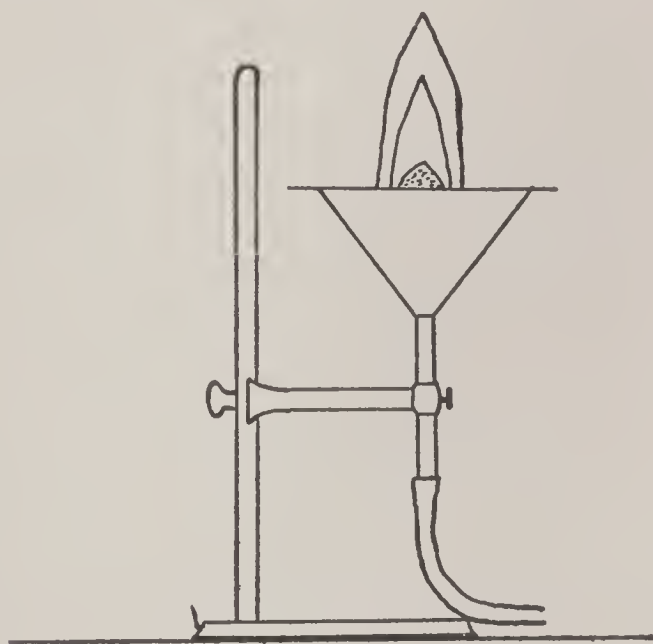


FIGURE 23.

Gunpowder enveloped in flame but still not burning.

Make a wire support for the chimney and mount it in the bottle with the upper end projecting a little distance above the top. Fit the constriction in the chimney with the gauze disk. Fill the chimney two-thirds full of lumps of calcium carbide. Insert in the stopper the length of glass tubing and attach to it the rubber connection, gas tip and pinch cock. Now fill the bottle nearly full of water and open the pinch cock (Fig. 24). The water will slowly rise into the chim-

ney and coming in contact with the carbide will generate acetylene gas. Gradually the air will be driven from the chimney and the gas jet may be lighted. You will obtain a soft white light of great brilliancy.

Smoke Rings.—The production of a gas that will

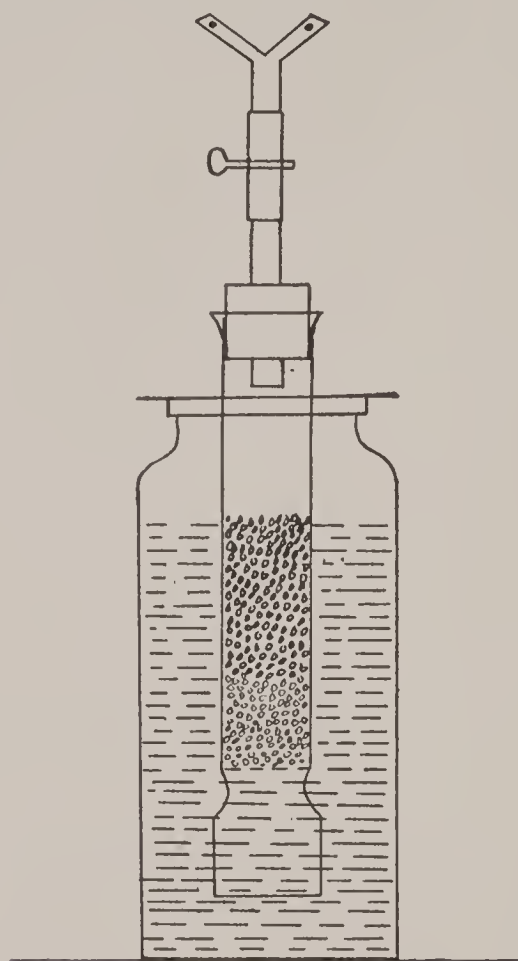


FIGURE 24.

An acetylene lamp.

undergo spontaneous combustion the instant it comes in contact with the air will make a very spectacular demonstration. Incidentally you will be able to obtain with it some of the prettiest smoke rings that were ever formed.

In a small flask put from 200 c.c. to 300 c.c. of a

strong solution of either sodium or potassium hydroxide. Into the neck of it fit two angle tubes, one reaching just through the stopper and the other extending nearly to the bottom of the flask. Mount the flask on a ring-stand over the Bunsen burner. Put into the solution a piece of yellow phosphorus about as large as a pea. (*Remember that phos-*

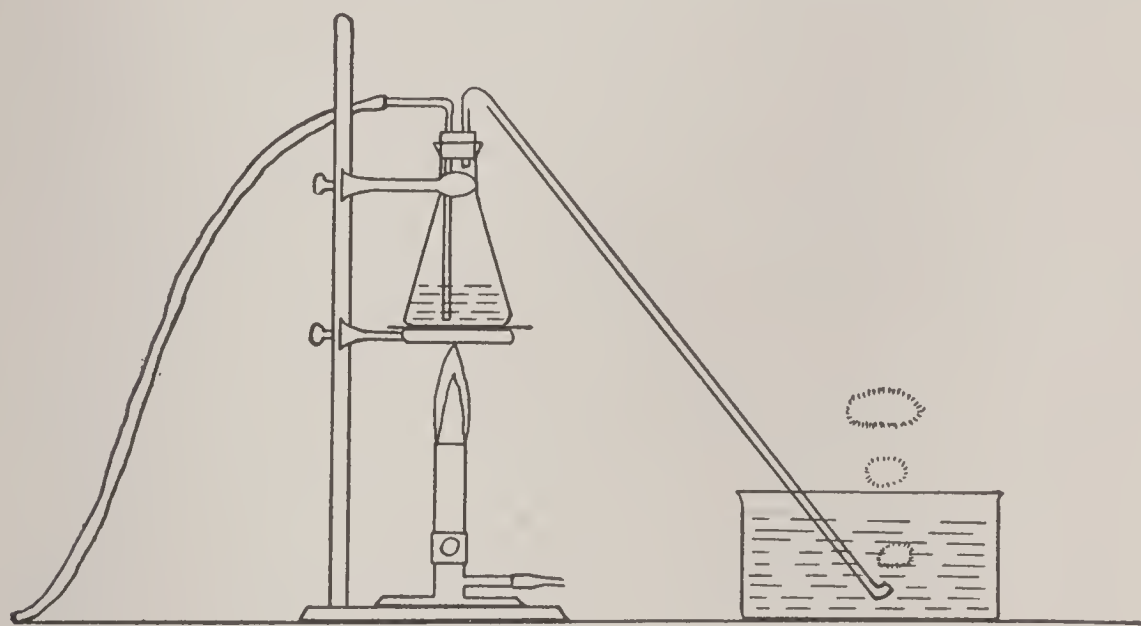


FIGURE 25.

Apparatus for producing smoke rings.

phorus must not be touched with the fingers and that it must be cut under water.) To the angle tube passing beneath the surface of the solution in the flask connect the supply of illuminating gas and to the other attach a delivery tube, the bent up end of which dips beneath the surface of water in a large basin. (Fig. 25.)

First open the gas cock and allow the illuminating gas to sweep the air from the flask. Then light the

Bunsen burner and bring the solution to a boil, keeping the end of the delivery tube under the surface of the water in the basin. When the solution is boiling vigorously, turn off the gas and watch the bubbles rise through the water in the basin. In a few moments they will spontaneously ignite and, as each bubble does so, a beautiful smoke ring will form and slowly rise in an ever expanding circle to the ceiling.

A beautiful green flame may be produced by the combustion of 20 c.c. of alcohol containing 3 g. of borax and 3 c.c. of concentrated sulfuric acid. Put the mixture into a flask provided with a one-hole stopper and short length of glass tubing. Mount the flask over the Bunsen burner and clamp just above the stopper a 2-inch length of combustion tubing. Boil the mixture in the flask and ignite the vapor that issues from the tube. It will burn with a green flame.

Fire Extinguishers.—One of the most difficult problems connected with combustion is to find efficient means of stopping the process when it gets beyond our control or starts by accident.

Into the bottom of a beaker put a teaspoonful of baking soda and pour upon it a little acid. Any acid may be used, or even vinegar. Into the gas that bubbles up and fills the beaker thrust a lighted match. It at once goes out. The carbon dioxide that envelops it will not support combustion.

Into an evaporating dish pour a little carbon tetrachloride, or "Pyrene." Heat it over the flame and thrust into the heavy vapor that forms a lighted

match. Again, robbed of oxygen, the flame goes out. Carbon dioxide and carbon tetrachloride are two of the most efficient agents used in fighting fire.

Put a teaspoonful of gasoline into an evaporating dish and light it with a match. Pour a little water on it, but note that the flame is not extinguished. The gasoline being lighter than the water floats upon its surface and continues to burn. Now pour a little carbon tetrachloride into the dish and the flame will quickly go out. This liquid vaporizes and surrounds the flame with a heavy blanket of incombustible gas and one which will not support combustion.

Making a Fire Extinguisher.—Fill a test tube two-thirds full of a saturated solution of baking soda. (A saturated solution is made by dissolving in a liquid all of the finely pulverized substance that it is possible to make it hold. It is often a good plan to stir a considerable quantity of the substance into the hot liquid and then cool the solution. If some of the substance separates out on cooling, the solution is saturated.)

In this solution float a small pill bottle containing a little concentrated sulfuric acid. Fit the neck of the test tube with a one-hole stopper carrying a glass elbow. Quickly invert the tube spilling the acid into the solution and immediately there will be a rapid rush of gas and froth through the exit tube. The gas is carbon dioxide, and, if the delivery tube is directed toward a lighted candle, the flame will be extinguished.

A larger fire extinguisher may be made in a similar way using a pint bottle. Fill the bottle two-thirds

full of bicarbonate solution (baking soda) and place a dilute solution of the acid in a test tube so resting in the bottle that its mouth is not submerged (Fig. 26). To the exit tube attach a rubber connection carrying a jet tube at the end. Leave a good sized opening at the end of the jet tube. Upon inverting this a vigorous flow of carbon dioxide will be obtained.



FIGURE 26.

A fire extinguisher. The bottle contains a saturated solution of baking soda and the test tube moderately strong sulfuric acid.

A Carbon Dioxide Trick.—Arrange a generator for producing carbon dioxide from marble and hydrochloric acid, just as you did in the chapter on the Atmosphere. Get the generator into action and let the delivery tube pass to the bottom of a wide mouth bottle placed upright on the table (Fig. 27). Since carbon dioxide is heavier than air, it will settle in the bottle and displace the air which originally filled it.

Have at hand a short candle to which you have attached a stout wire so bent that you can raise and lower the upright flame in the bottle. You can tell how full of carbon dioxide the bottle is at any time by lowering the lighted candle into it. The flame will be ex-

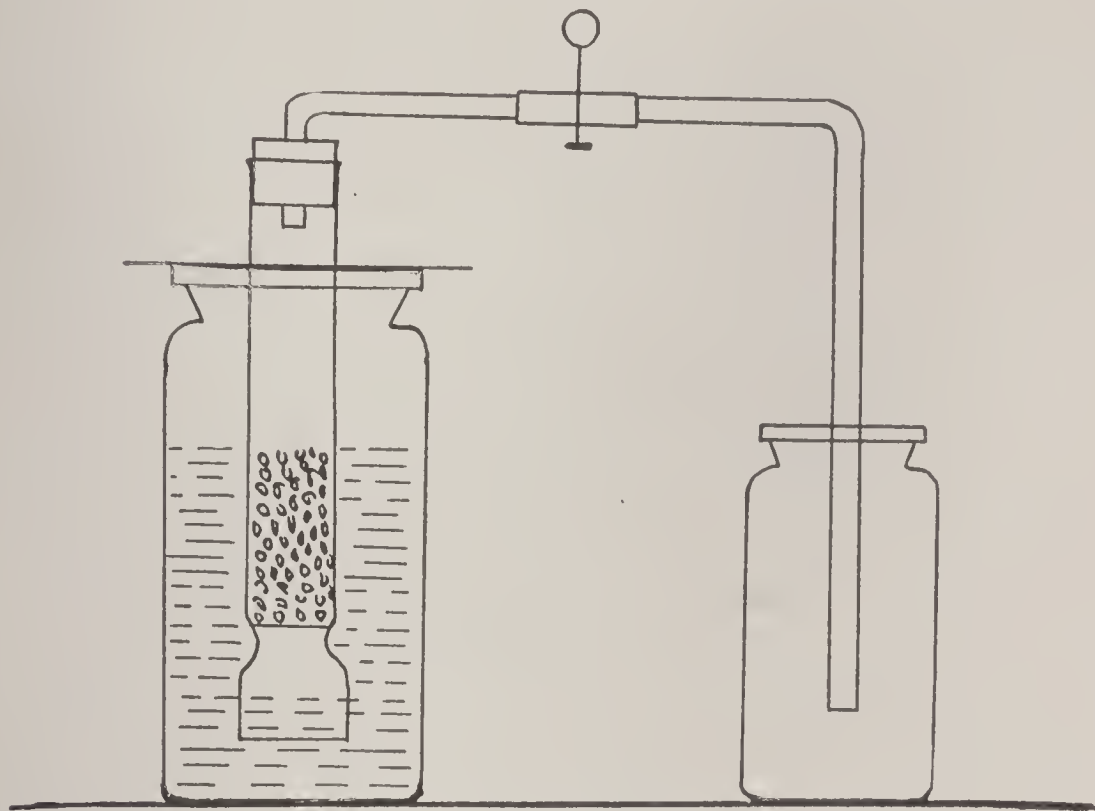


FIGURE 27.

Convenient form of apparatus for the preparation and collection of carbon dioxide. The chimney contains marble chips resting on a perforated lead disk, and in the bottle is dilute hydrochloric acid.

tinguished at the level of the gas. And here is the trick. If, just as the flame begins to go out and apparently does so, you will quickly raise the candle the flame will reappear. A little practice will enable you to do this beautifully. The explanation is this. A tiny almost colorless flame still burns just at the level of the carbon dioxide and is connected with the wick

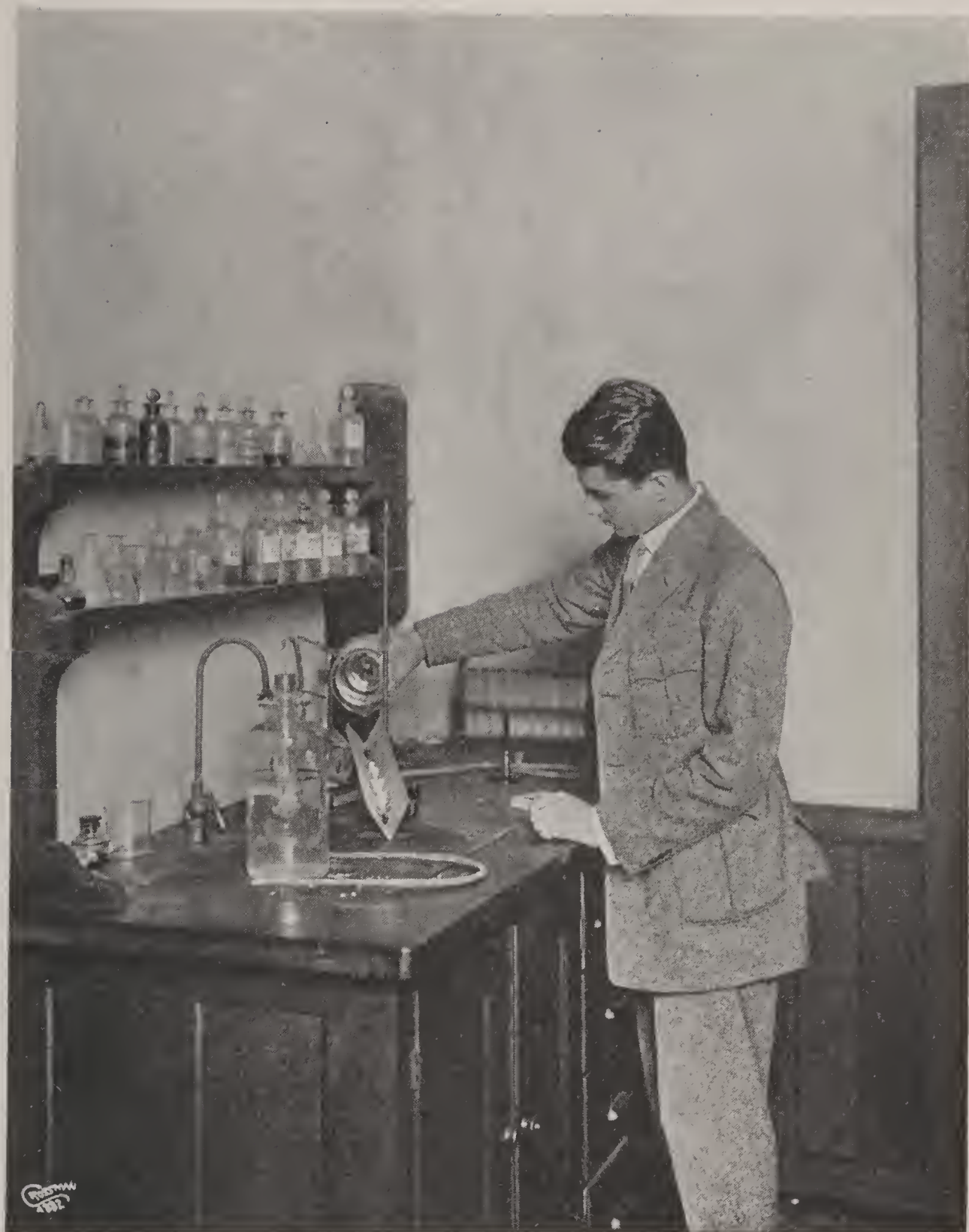
by an invisible column of still hot incandescent vapor. Upon quickly raising the candle this vapor burns and relights the wick.

A more striking demonstration of the fire extinguishing properties of carbon dioxide may be shown in the following way. Make a small trench about 18 inches long and 3 inches deep by folding a piece of sheet zinc trough shaped. Set in the bottom of this a row of Christmas candles. Soften the bottoms of the candles and they will adhere to the zinc upon cooling.

Start your generator and fill a quart bottle with carbon dioxide by displacement of air, as before. Raising one end of the trench on a small block, light the candles and quickly pour down it the bottle of gas. One after another the candles will be extinguished. This interesting experiment illustrates several properties of this gas. It is heavier than air. It does not support combustion. It is incombustible.

Carbon monoxide, an important constituent of a number of fuel and illuminating gases, may be prepared by the action of concentrated sulfuric acid on formic acid. (*Since this gas is poisonous be careful not to breathe it.*)

The apparatus for the preparation of this gas is exactly the same as that for hydrogen. Pour through the thistle tube a little formic acid. Be sure that the thistle tube dips beneath the surface of the liquid in the generator. Then add to the contents of the generator some concentrated sulfuric acid. A gas will



Extinguishing candles by pouring carbon dioxide upon them.

at once begin to escape which may be collected over water.

Remove a bottle of the gas and present the mouth of the bottle to the Bunsen flame. A blue flame will fill the bottle showing the combustible nature of the gas. Shake a little limewater in the bottle and the white precipitate will show that carbon dioxide is the product of the combustion. This is the gas that burns with a blue flame when you open the furnace door shortly after fresh coal has been put on the fire. It is also the poisonous gas which issues from the exhaust of an automobile engine when it is run in a closed garage where the supply of air is insufficient for complete combustion of the gas.

Producer gas, one of the most important fuel and power gases, consists principally of carbon monoxide. If you have made an electric furnace of the combustion type, fill the silica tube with small pieces of charcoal and connect to one end of it a carbon dioxide generator. To the other end connect a delivery tube and arrange a pneumatic trough with inverted bottles of water. Bring the charcoal to a good heat and pass over it a *slow* stream of carbon dioxide. To do this use in the generator dilute acid and only a little at a time. Collect in the bottles the gas that escapes from the tube.

Remove one of the bottles and apply a match to the gas. It burns with a blue flame showing that the carbon dioxide has been reduced to the monoxide. This is essentially what happens in a gas producer. Like-

wise in a furnace carbon dioxide forms in the bottom of the fire-pot and is reduced by the fresh coal on top.

Carbon monoxide is the gas which often kills people in poorly ventilated sleeping rooms. Fresh coal is put on the fire at night and the dampers nearly closed. Possibly, too, the lids are partly removed to keep the fire from burning out. Here are just the conditions for the production of this deadly gas, which is all the more dangerous because it is odorless.

The Flashing Point of Kerosene.—Place 2 or 3 tablespoonfuls of kerosene in an evaporating dish and *remove the kerosene container to a safe distance*. Try to light the kerosene with a match. If it is of good quality, you will not succeed. Very gently warm the mixture with a small flame and at frequent intervals apply a match to the vapor. With a thermometer take the temperature of the liquid and note the point at which the vapor just ignites. This is the *flashing point* or very nearly it. If the vapor continues to burn, you have passed the point. In that case put out the burner flame and allow the kerosene to cool. Every minute or two test with a lighted match. The temperature at which the vapor will just light is the flashing point. It should lie between 120° F. and 140° F. This point is the most important indication as to the quality of this liquid fuel.

Fireproofing Cloth.—Make a solution of 16 g. of alum, 3 g. of boric acid, 5 g. of ammonium carbonate, and 3.5 g. of borax in 100 c.c. of water. Dip into it a strip of muslin and let it dry completely. Try to

burn it, and compare the rate at which it burns with that of muslin not so treated.

Making a Safety Match.—Make a mixture of 4 drops of thin glue, one-fourth as much red phosphorus, and as much finely pulverized sand as phosphorus. Spread the mixture in a thin layer on a piece of pasteboard, and allow it to dry for a day.

In the meantime whittle out some match sticks. Melt some paraffin and soak the ends in it. Make a mixture of 3 drops of *thin* glue, twice as much powdered potassium chlorate, and one-third as much powdered antimony sulfide. Dip the sticks into the mixture and then let them dry for several hours. When they are dry, you will be able to light them by rubbing on the prepared surface.

In this chapter we have described only a few of the many experiments on combustion which may be performed, but they are sufficient to give you a fairly broad understanding of this important chemical process.

CHAPTER 22

JOSEPH HENRY

NOT all of the distinguished scientists who started as home-laboratory workers became chemists. One of the earliest and most eminent of America's scientific men was the physicist Joseph Henry. He was born near Albany, New York, in the year 1799. His first ambition was to become an actor, but one day he chanced to read a book on science, which he ever after said changed the whole course of his career. He at once determined to devote his life to scientific work. To obtain the necessary education Henry attended the Albany Academy, and, to pay his expenses, he taught school at intervals. After a time he was engaged to survey the route for a new highway through the untracked forests of the Empire State. He did it during the dead of winter, amid deep snows and often exposed to the danger from wild beasts. So interesting did this work seem to him that he was about to accept another engineering contract when he was offered the position as professor of physics at the Albany Academy. This decided his future, and fortunate it was for the cause of science in America.

Henry plunged into his new duties with enthusiasm, and immediately he began to experiment in every spare moment. Much of his work was original, and

all of it was new to him. Particularly did he distinguish himself in the field of electricity. Independently of any knowledge of Faraday's discoveries with induced currents and anticipating him in many of them Henry covered the same ground. Sturgeon in England had just invented the electromagnet and Henry determined to try out the possibilities of this new device. Going to a blacksmith-shop he obtained a rod of soft iron bent into the shape of a horseshoe. Not finding any wire suitable to his purpose, he invented silk-covered magnet wire. Then he wound upon this soft iron core 9 lengths, each 60 feet long, of fine insulated copper wire, arranged so that he could connect them in any desired manner. With this little giant and a cell containing only a half pint of acid and one-fifth of a square foot of metal surface, Henry lifted masses of iron weighing from 60 to 650 pounds. With a tiny cell having plates only one inch square, he lifted 85 pounds. For Yale and Princeton universities he constructed large magnets capable of lifting considerably more than a ton. Henry also invented the relay, without which telegraphy would have been impossible. Through a mile of wire strung back and forth across his laboratory he actually sent and received signals.

Later Henry became professor of physics at Princeton University and was made first secretary of the Smithsonian Institution in Washington. And all of this and much more came from a boy's early desire to experiment.

CHAPTER 23

FIREWORKS

I PRESUME no other part of chemistry appeals to the beginner more strongly than spectacular displays of fireworks. You will wish to celebrate the Fourth of July with plenty of light and noise, and yet do so without danger to yourself. I will try to show you how this may be done.

Touch Paper.—In igniting some of the mixtures which I shall describe you will need fuses. Touch paper is ideal for this purpose. It is prepared by soaking strips of filter or blotting paper in a strong solution of potassium nitrate (saltpeter) and allowing them to dry. This paper will burn without a flame, just as a fuse does, and the fire cannot be extinguished by blowing.

White Fire.—Mix together equal quantities of powdered potassium chlorate and magnesium dust. Use about two teaspoonfuls of each. Place the mixture on a square of asbestos board, or a block covered with asbestos paper, and *ignite it with a long wax taper. Never use a match.* The combustion comes with great suddenness and gives a flash of blinding light and a dense cloud of smoke. At night it is of wonderful brilliancy. This mixture is commonly called flash

powder, so much used in taking indoor photographs. The light is particularly rich in the chemical rays that affect a photographic plate.

Red Fire.—Mix 1 g. of powdered potassium chlorate with 11 g. of strontium nitrate. If these substances are not already powdered, pulverize them *separately* and then mix them. *Never pulverize them together.* Stir into this mixture 4 g. of finely powdered sulfur and $\frac{1}{2}$ g. of lampblack. Make a cone-shaped heap of the mixture on an asbestos square and insert in the top a piece of fuse paper. Ignite the fuse and the mixture will burn with an intensely red flame.

Green Fire.—This time mix 3 g. of pulverized potassium chlorate with 8 g. of powdered barium nitrate and 3 g. of powdered sulfur. Again do not pulverize the chlorate and nitrate together. Using a fuse of touch paper ignite the mixture. It will give a brilliant green light.

Purple Fire.—*Separately* pulverize 2 g. of copper sulfate (blue vitriol), $2\frac{1}{2}$ g. of sulfur, and 15 g. of potassium chlorate. When ignited with touch paper, a purple fire is produced.

Blue fire may also be obtained from a mixture of 2 g. of pulverized charcoal, 2 g. of cupric chloride, and 4 g. of potassium chlorate.

The National Colors.—Try at night the ignition of a mixture of the materials used in red fire, white fire, and blue fire.

Igniting flash powder by electricity is an interesting and safe method to use. Cover a board 6 inches

square with asbestos paper. Along the middle line about 3 inches apart drive two nails, leaving an inch of each above the board. If you like, you may insert brass binding posts instead of the nails. Between the nails connect a piece of No. 30 iron wire, or any other piece of very fine iron wire which you may have. Over the wire place in a heap the mixture of powder which you wish to ignite. Off at some distance, 4 or 5 feet, place a half dozen good dry cells. Connect them through a switch to the ends of the iron wire on the ignition block. Then close the circuit. The electric current will heat the iron wire to incandescence and set off the powder.

Old-fashioned Gunpowder.—Before the invention of the powerful explosives of modern times, gunpowder consisted of a mixture of charcoal, sulfur, and saltpeter. Here were two easily combustible substances and a strong oxidizing agent. Furthermore the products of the combustion were chiefly gases and a large quantity of heat, just the combination needed for an explosive.

To prepare this powder make a mixture of 30 g. of finely pulverized potassium nitrate and 5 g. each of powdered charcoal and sulfur. Ignite it with a long wax taper. It will burn readily giving a dense cloud of smoke, and the unavoidable smoke was the great objection to this kind of powder. In addition, the oxygen necessary for the combustion must come from particles outside the charcoal and sulfur. In modern high explosives the molecules of the explosive com-



Throwing the switch for the ignition of flash powder. A 110-volt circuit with no resistance is being used.

pound contain the oxygen necessary for their own combustion. This results in a much quicker and a vastly more violent explosion.

Making a Sparkler.—In a small test tube melt some potassium chlorate. When it has cooled to a solid mass, break away the glass. Rub this stick on the ignition surface of a safety match box. A shower of sparks will result.

Combustion of Sugar and Potassium Chlorate.—Carefully mix without friction equal quantities of finely powdered potassium chlorate and granulated sugar. Place the mixture on the asbestos square and by means of tongs drop upon it a piece of asbestos paper saturated with the strongest sulfuric acid. A very rapid combustion and an intensely blue flame will result.

Combustion of Aluminum Dust and Bromine.—Clamp a test tube in a vertical position on your ring-stand and place in it a small quantity of aluminum dust or filings. Heat the bottom of the tube until the glass is just red. Then from a medicine dropper let fall upon it 10 drops of bromine. A very vivid combustion results.

After the contents of the tube has cooled let 2 or 3 drops of water fall into it. You will obtain another vigorous action and a hissing sound. (For this experiment a hard glass test tube is best.)

Combustion of Aluminum Dust and Iodine.—The previous experiment may be repeated with iodine, which will be more easily obtained. Heat the alumi-

num to faint redness in a hard glass test tube and drop upon it from a folded paper a half gram of finely powdered crystals of iodine. The combustion is even more vivid than it is with bromine, and, if a Bunsen burner flame is quickly held to the mouth of the tube, the product of the combustion itself will take fire.

Combustion of Zinc Dust.—Mix 8 parts of ammonium nitrate with 1 part of ammonium chloride. Spread this out in a thin layer on a plate of glass and cover it with a layer of zinc dust. Let fall upon the mixture a single drop of water and combustion will take place.

A Gas Explosion.—But I hear you asking, “Where is the noise?” A Fourth of July celebration would surely be a fizzle without an explosion. The following perfectly harmless demonstration will provide you noise in abundance. Obtain a gallon can having one hole in the top and cut another. Fit one hole with a cork carrying a short glass elbow and connected by rubber tubing to a supply of gas. Through the other hole insert a cork carrying a straight glass tube of about $\frac{3}{8}$ of an inch inside diameter and extending into the can about 8 inches. If this second cork does not fit tightly, seal it in with sealing wax. Then turn on the gas and in a moment or two light it at the top of the straight tube. When the gas is burning up well, remove the cork carrying the elbow and as you do so turn off the gas. The flame will continue to burn at the top of the straight tube, but it will grow smaller and smaller and finally almost disappear (Fig. 28).



Waiting for the flame to strike down and ignite the explosive mixture. Using a 2-necked bottle instead of a can.

You may think it has gone out, but it hasn't. It is still playing about the mouth of the tube or reaching slowly down to the explosive mixture accumulating inside the can, for, as the gas escapes and burns, air enters through the open vent. Presently the flame

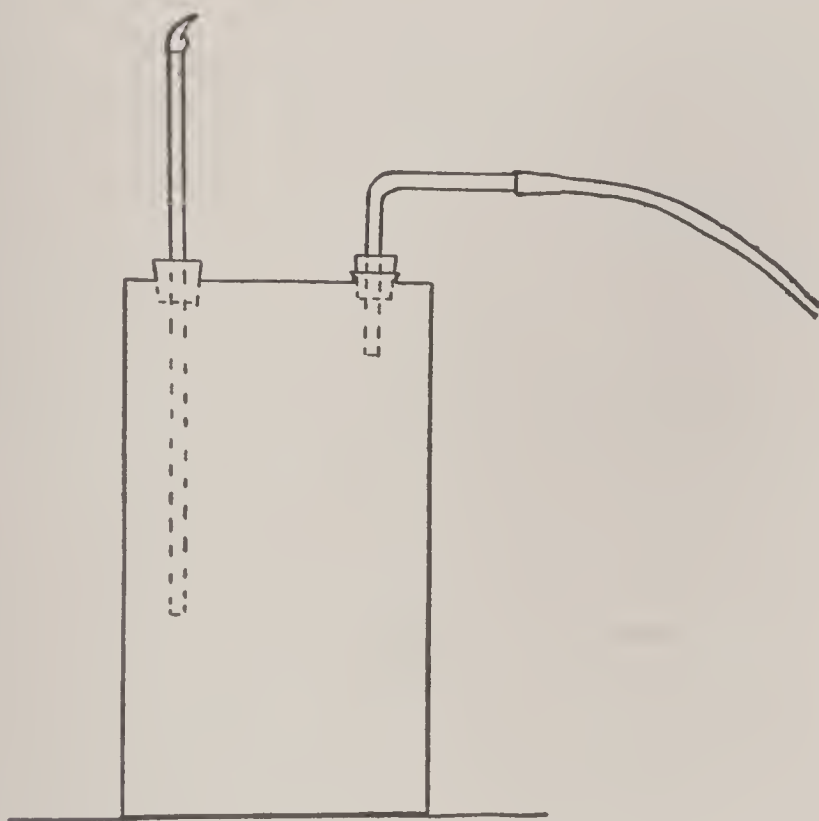


FIGURE 28.

The set-up for a gas explosion. As the can fills with gas, a match is applied to the tube at the left and at the same time the tube at the right is withdrawn from the can.

strikes down igniting this mixture and producing an explosion that will satisfy every requirement of a Fourth of July celebration.

This experiment also illustrates the striking back of a Bunsen burner or a gas stove and the back-firing of an automobile engine.

Another Explosion.—Into this same can but with the corks removed pour from 5 to 10 c.c. of gasoline or ether. *Keep both away from the flame.* Wait about 2 minutes and then with a long taper bring a flame over one of the openings. The explosion will be all that you desire. *In lighting this explosive mixture never put your face over the can.*

Nitrogen Iodide.—Now I am going to tell you how to prepare an explosive substance that will afford you no end of amusement. But you must be careful and do exactly what I tell you to do. From the drug store obtain a few grams of iodine crystals—not tincture of iodine. In a test tube place 4 or 5 c.c. of alcohol (denatured will serve) and add a considerable quantity of the powdered iodine. Shake this so as to obtain a very strong solution. Mix with it an equal volume of the strongest ammonia water. Shake the mixture. A black precipitate of nitrogen iodide immediately separates. Pour this onto a filter paper placed in your funnel taking pains to get out all of the solid. If any remains in the test tube pour the filtrate back into the tube and, after shaking, quickly pour it onto the filter.

Now this substance is not at all explosive when wet and it will take about an hour for it to dry, possibly longer. Remove the filter paper, tear it into several pieces, and distribute them in a number of out-of-the-way places. *Then do not again touch them with your hands or go near to them.* After about an hour tickle one of the pieces with a feather attached to a rod 3 or

4 feet long. Even this slight friction will cause the nitrogen iodide to explode. And the noise will be appropriate to the occasion. A jar like the stamping of a foot will cause it to explode. Blow upon another portion of the stuff through a long glass tube, and the pent-up energy will be released. If another portion has been left in the sun or in a warm place it will explode spontaneously upon the slightest provocation.

A chemistry "prof" of my acquaintance left some of this treacherous stuff in various places about his lecture room to dry one night. During the evening the janitor entered the room, and, as he shut the door, a sharp explosion startled him nearly out of his wits. Then, as he began hurriedly to walk through the room, explosion after explosion greeted his progress. Alarmed and terrified, he made his exit and, doubtless to this day, congratulates himself upon his narrow escape.

Wet nitrogen iodide cannot explode, and, after setting it to dry, simply keep at the proper distance. When the psychological moment arrives, it will know perfectly well how to behave.

In the foregoing experiments I have given you plenty of ammunition for an old-fashioned Fourth of July celebration. There are other experiments of this sort, but many of them are dangerous unless carried out with the utmost care.

CHAPTER 24

SIR HENRY BESSEMER

As a lad Sir Henry Bessemer, who later discovered the quick process of blowing molten pig iron into liquid steel, displayed a wonderful genius for the use of tools and the work of invention. Before he was out of his teens, he had devised a method for stamping deeds, a machine for figuring velvet, and a type-casting device. Then one day he had occasion to buy some bronze powder and to his amazement found that it cost a fabulous price. Why should it cost so much, he asked? If he could discover the secret of its preparation he felt that he could make his fortune. Immediately he set to work. He designed a machine for turning bronze into what seemed to him a powder of the utmost fineness. But still it was not fine enough. Baffled, he abandoned the project for a time. Then someone suggested that he examine his powder with the microscope. He did so and at once the difficulty was revealed. The particles of his powder were little curled-up shavings. This gave him his clew and he quickly designed a machine which accomplished his purpose. But he never patented it, fearing that someone would steal his process. Under the supervision of men sworn to secrecy, Bessemer manufactured the

powder and sold it at a huge profit. For years this source of revenue supplied him with all the funds he needed for his work of experimentation and invention.

And to invent was as natural to Bessemer as it was to breathe. In a short time he turned his attention to the making of guns and devising improvements in their designs. This work required steel, and Bessemer soon became convinced that a cheaper and quicker process of producing this vastly important raw material was imperative. With him always the realization of an industrial need was a summons to meet it. Long he pondered the problem. Then one day a happy thought came to him. Why not burn the impurities out of molten pig iron and convert it into steel with a blast of air blown through the iron placed in an egg-shaped crucible. He tried the plan on a small scale in his laboratory and it worked. On a bigger scale, it was at first a failure, and Bessemer became the laughing stock of England. But the old proverb is right, "He who laughs last laughs best." Bessemer perfected his process and he was soon manufacturing steel in his own works at Sheffield and putting it upon the market at \$100 a ton cheaper than his competitors could produce it. That was an argument that could not be met, and the very men who had ridiculed him were soon paying him a fortune in royalties for the use of his process.

For his services as an inventor Bessemer was knighted. His remarkable success in everything he undertook grew out of his boyhood ambition to experiment.

CHAPTER 25

THE CHEMISTRY OF THE ELECTRIC CURRENT

TO the amateur there is probably no more fascinating branch of science work than that of electricity. There is a magic about this word that always appeals to boys. One of the first things that my General Science boys want to do is to electroplate something. And I presume you are no different than they. Let us see what can be done.

Source of Current.—The first all-important matter to consider is where you will get the supply of current, or “juice,” as the boys say. If you have in your house direct current and have made the lampboard rheostat that I described in the chapter on electric furnaces, the problem is easily solved. Simply run an extension coil from a lamp socket to the lampboard and connect through that with your apparatus. For much of this work you will not need an elaborate lampboard. Just mount 2 or 3 lamps in parallel (in the same manner as the lamps on the board diagramed in a previous chapter) and turn on one or all, as you may need. A single lamp will frequently be sufficient.

You who are not so fortunate as to have current in

the house or who have alternating instead of direct must make cells and batteries. But you can do that.

The bichromate cell will be best for work in which you will need much voltage or considerable current. A half dozen of these coupled in series will meet most of your needs, aside from those of electric furnaces. To make such a cell obtain a quart fruit jar and place in it 80 g. of chromic acid. Pour onto this 710 c. c. of water. Stir the mixture until the acid is thoroughly dissolved. Then add slowly and with constant stirring 45 c. c. of concentrated sulfuric acid. Now to obtain electricity by chemical action you must insert in this solution a rod of zinc and a rod of carbon. A broad flat bar of zinc about a quarter of an inch thick will be best for the negative electrode. Solder to it a copper wire for connections. For the positive electrode a large-size electric light carbon will serve, or better the carbon rod from an old dry cell. By means of a wooden clamp holder, which you can devise yourself, support these in the jar. Do not let them touch and see that they dip into the solution as far as possible. If you have used the carbon from a dry cell, it will have a binding post. If not, you may twist about the top of it with pliers a stout copper wire. When not in use remove the zinc from the solution.

To connect these cells in series join the carbon of one to the zinc of the next with short lengths of copper wire. That will leave a carbon of one end cell to be connected through your apparatus with the zinc of the

other end cell. A half dozen of these cells will give you 12 volts.

The gravity cell is best adapted to electrotyping and electroplating. For such work you do not need a large current but a constant one and a cell that can be used on closed circuit. To make this cell obtain a half gallon battery jar. Suspend in it a large square of

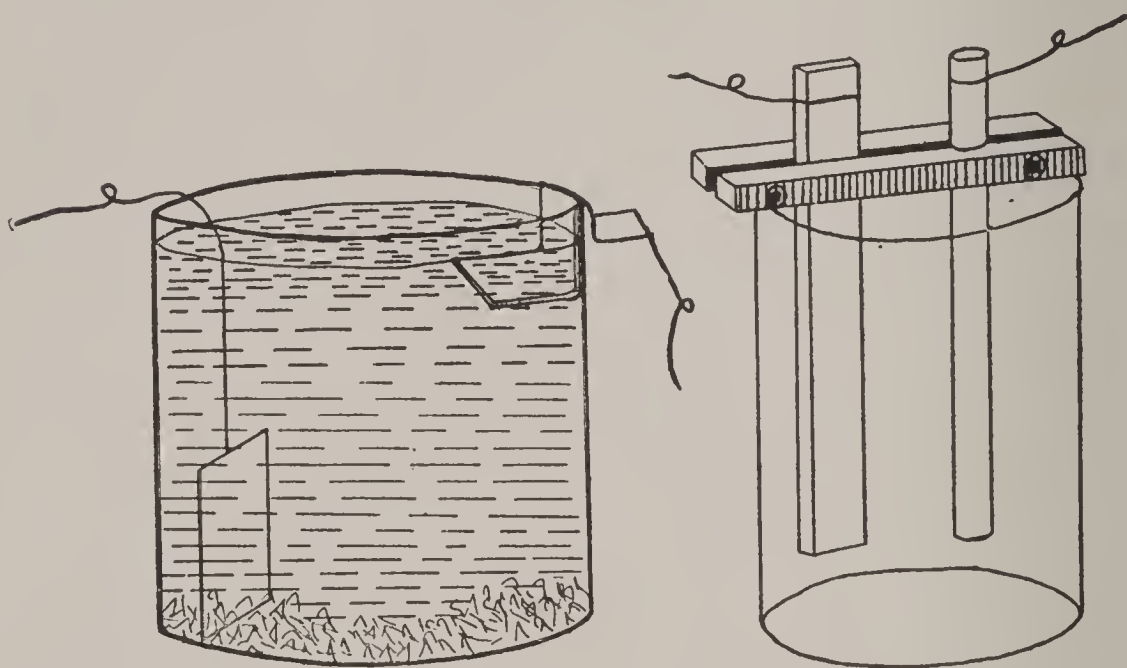


FIGURE 29.

Gravity cell at the left, and electrodes and holder for bichromate cell at the right.

sheet copper letting the lower edge rest upon the bottom and the upper edge extend about half-way up the jar. Over the edge of the jar suspend a heavy bar of zinc bent hook-shape at one end to catch on the edge of the jar and extending horizontally inside the jar. To both of these solder copper lead wires. (Fig. 29.)

Into the bottom of the jar throw a handful of blue

vitriol and fill it with water containing a few cubic centimeters of sulfuric acid. Then place the cell on closed circuit. It should never be left on open circuit. The blue vitriol will gradually be used up and more must be added from time to time. The zinc also must occasionally be replaced.

Dry cells will serve many purposes, but they are not adapted to closed circuit work, such as electroplating. This is because they polarize rapidly and the current becomes very small. Where you need a large current for a short time, six good dry cells will be just the thing.

A rod of carbon and a rod of zinc placed in a jar containing a strong solution of sal ammoniac will give you a cell similar in principle to a dry cell and suitable for many purposes.

Electrolysis Apparatus.—If you are to carry out the electrolysis of water and solutions, some form of electrolysis apparatus will be required. It will be practically as cheap and much more satisfactory to buy a simple form of such apparatus from some supply company than to attempt to make one yourself. To make one will require platinum foil, and that is quite expensive.

For the boy who wishes to do so, though, I am going to describe the construction of a simple form. Obtain two small strips of thin platinum foil. They need not be more than a quarter of an inch wide and three-quarters of an inch long. Pierce each with a small hole and attach to it a stout length of copper wire well

insulated with heavy rubber. Bend both ends hook-shape and suspend the two electrodes in a small battery jar containing water to which has been added a little sulfuric acid. Fill two test tubes with the same solu-

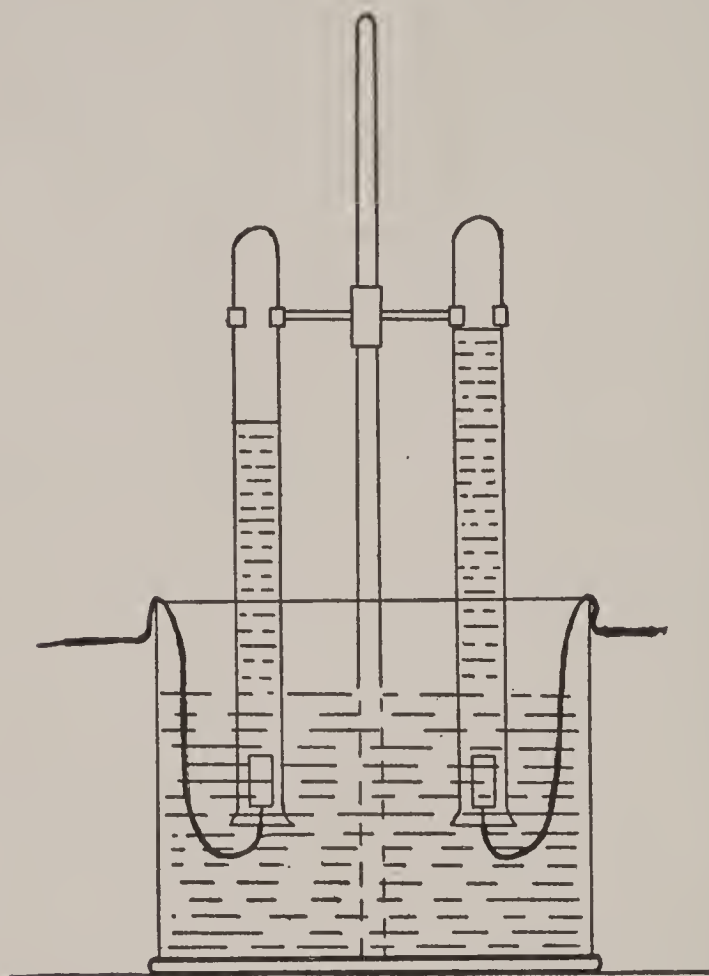


FIGURE 30.

Electrolysis apparatus.

tion and placing a thumb over the mouth of each invert them in the jar. Slip their mouths over the platinum electrodes and secure them in position by clamps attached to your ring-stand. (Fig. 30.)

If you have direct current, connect the lead wires of the electrolysis apparatus with one lamp of the rheo-

stat. If you depend upon cells, connect a half dozen bichromate cells directly with the apparatus. In either case you will immediately obtain action. Bubbles of gas will rise from each electrode and collect in the test tubes. You will note that the gas accumulates twice as fast in one as in the other. This is from the negative pole, or *cathode*. The other electrode is called the *anode*.

When the cathode tube is about two-thirds full, break the circuit. Loosen the clamp and placing your thumb beneath the mouth of the test tube containing the smaller quantity of gas quickly invert it and thrust into it a glowing splint. The brilliancy with which the gas supports combustion proves it to be oxygen. Simply lift the other tube from the water and present it to the flame. The explosion which you get together with the mist which collects on the walls of the tube shows this gas to be hydrogen.

The other type of apparatus, which you can obtain from a supply house, is simpler and requires a smaller quantity of solution. If you use it, the openings at the tops of the tubes will need to be closed with short pieces of rubber tubing and pinch clamps. To use this apparatus open the pinch clamps and fill the tubes to the top with the solution. Then connect the electrodes with the source of current. Close the pinch clamps and allow the gas to collect. In testing for oxygen, hold over the tube a glowing splint and slowly open the clamp. The splint will be kindled into a flame. To obtain the hydrogen test, hold an in-

verted test tube over the other tube and open the clamp. Being lighter than air the hydrogen will rise and fill the tube. Hold the tube of gas to the flame and it will burn with a sharp explosion.

Testing Electrolytes and Non-Electrolytes.—Some substances in solution will conduct the electric current and others will not. Those which do are called *electrolytes* and those which do not are called *non-electrolytes*. It will afford you an interesting piece of real research to test out with your apparatus the electrical conductivity of a number of compounds, some organic and some inorganic. Since chlorine attacks platinum, it will be impossible for you to use the solution of a chloride, as common table salt for instance.

First rinse out the apparatus thoroughly and do so again after each trial. Then try the action of the current on distilled water. Is there any action? Is distilled water a conductor? To determine this observe whether any bubbles appear at either electrode. If they do, the substance is an electrolyte. The bubbles may be very small and very few. If so, the substance is a very poor conductor. If you are using a 110-volt source of direct current and a lamp for resistance, the lamp will light, if the solution is a good conductor. I think you will find that distilled water is a non-conductor.

Next try tap water. You will probably get a slight action in this case, because tap water always contains some mineral matter in solution. The extent to which it conducts the current is an indication of the quantity

of mineral matter it contains. You have already seen that water containing sulfuric acid is an electrolyte. In order that water may conduct the current, it is necessary that it have some substance in solution, and the compound in solution is really the electrolyte. But not all compounds will make water a conductor.

Now place in your apparatus a quite strong solution of sodium sulfate which you have colored with blue litmus. Pass the current for a few minutes and you will note that the solution about the cathode becomes a deeper blue, while that about the anode changes to red. You will remember that an acid turns blue litmus red, while a base gives the reverse change. Chemical changes are being produced with the formation of an acid at one pole and a base at the other. *Electrolysis* is the name given to the chemical changes which occur when a current is passed through the solution of an electrolyte.

Test the conductivity of alcohol and of a solution of sugar. You will be unable to detect any action. These are organic, or carbon, compounds, and most organic substances are non-electrolytes. That is one of their chief distinctions from mineral compounds.

The strength of an acid or base may be determined by its ability to conduct the electric current. Dissolve 4 g. of sodium hydroxide and 5.6 g. of potassium hydroxide in separate beakers, each containing 100 c.c. of water. Test the conductivity of each solution separately, rinsing the apparatus thoroughly after each test. Which seems to be the better conductor and the strong-

er base? Test also the conductivity of ammonia water and limewater. Can you understand now why ammonia water may be applied directly to the clothing in the removal of stains, and why limewater may be used as an eye wash and even taken internally?

Place some vinegar in your apparatus and pass the current. How does its conductivity compare with that of sulfuric acid? Do you understand why vinegar is used on our food? Prepare a saturated solution of carbon dioxide by passing the gas through water for some time. Then determine whether its solution is an electrolyte. Why can you drink without injury such "vast" quantities of soda water?

Electroplating.—Now we are coming to what you have been waiting for. To have been able to electroplate another substance with a bright shiny metal would have intoxicated the alchemist with joy. But the electric current was unknown in his day.

Place 250 g. of blue vitriol (copper sulfate) in a battery jar and cover it with a liter of water (1000 c. c.). Stir it well and allow the mixture to stand overnight. Thrust into the solution a carbon rod from an old dry cell and a strip of copper having about the same square surface as that of the carbon rod. Using either two bichromate cells or two dry cells grouped in series, that is, the positive pole of one joined to the negative pole of the other, connect them with copper wire to the electrodes of the electroplating bath. Connect the positive pole, or the carbon, of the battery to the copper strip and the negative pole, or zinc, of the



Electroplating with current from gravity cell at the left. Testing electrolytes at the right. Lighted lamp in series with apparatus shows that the solution is an electrolyte.

battery to the carbon rod. Allow the current to pass for 5 minutes and then examine the carbon rod. Note the deposit of bright copper. Also examine the copper strip. Does it look any different than before you passed the current?

Now reverse the connections between the battery and the electrodes of the bath. Pass the current again for 5 minutes. Then examine the carbon rod. Is it still bright? What has happened to the copper deposit which you observed before? Examine the copper strip. Does it look any brighter? Do you not see that copper is driven out of solution by the current and deposited upon the negative pole of the bath, and that at the same time copper is dissolved from the positive pole of the bath? For these reasons we place the object to be plated at the negative pole, or cathode, and make the positive pole, or anode, of copper. In this way the strength of the solution is kept the same, but the copper anode is gradually used up.

If you use direct current and a lamp board rheostat, the process is exactly the same. But you will have to determine which wire is positive and which negative. To do so turn on the current and dip the ends of the copper lead wires into the copper sulfate solution for a few moments. The one upon which you obtain a copper deposit is the negative. For this work use a single lamp.

To do a good piece of electroplating work a battery consisting of two gravity cells grouped in series will be best. It will give a small but steady deposit, and that

is what you need for a fine even coating. Clean a nail or an old knife blade thoroughly with fine sandpaper. Scour it until it is bright. Hang it at the cathode of the electroplating bath, that is, you will connect to it the zinc pole of the battery. For the anode use a strip of copper as before. Allow the current to run a considerable time. You can tell when to break the circuit by the amount of the deposit. You should obtain a deposit of good quality. Whenever the deposit is dark and spongy and does not stick well to the surface, you are using too much current and the surface was probably not well cleaned. The current will not be too large with gravity cells. When you use direct current without proper resistance, it will be.

Making an Electrotpe.—An electrotpe is a copper plate from which a page of a book is printed or some "half-tone" or "line-cut" is reproduced. Phonograph records are also stamped from electrotpes.

It will be interesting to make an electrotpe. In a small shallow basin melt lumps of beeswax, making a layer at least an eighth of an inch thick. When the wax has hardened, rub it over thoroughly with finely powdered graphite. Then select some object you wish to electrotpe. It may be any metallic souvenir having a fairly smooth surface with or without inscription. But never attempt to copy a coin. That is illegal. Dust the object with graphite and press it firmly into the wax until a deep clean-cut impression is made. With a knife cut out and remove the square of wax containing this impression. Now dust the wax again being sure

that the graphite is worked into all of the lines and grooves. Rub it to a smooth shiny surface. Finally attach a copper wire to the upper side of the wax form and suspend it at the cathode of the copper plating bath. At the anode, as before, suspend a strip of copper of the same area as that of the wax. Connect the bath with two gravity cells and allow the current to run for a considerable period. Probably 24 hours will not be too long, but the time will depend upon the thickness of the layer of copper that is deposited.

To make the electrotype permanent it will require backing with melted tin. When the required thickness of copper has been obtained, remove the form from the bath and carefully pry the electrotype loose with a knife blade. With molding clay make a shallow form of the exact shape of the electrotype. Have the bottom perfectly smooth and level. Place the electrotype face down in this form being careful to make the edges fit close up to the sides. Put this form in an oven and let it come up to as high a temperature as possible. Melt a quantity of tin in an iron dish and holding the dish with tongs pour the molten metal into the form. Then let the oven and contents cool. Remove the electrotype, ink it, and take off a copy on white paper.

Alundum cement or fire clay and water glass will be excellent material for making the form.

Silver Plating.—If you wish to silver plate, the process is exactly the same as that of copper plating. In a small battery jar place a dilute solution of silver nitrate. If you have prepared silver nitrate crystals

as directed in the chapter on metals, they will be suitable for use here. One gram to 100 c. c. of water will be sufficient. For an anode this time you will need a silver piece. Attach a wire to it and suspend it as you did the copper strip. The object which you plate must be perfectly clean. Scour it with sapolio and fine sand and rinse in clean water. Use gravity cells, as before, and allow the current to flow for some time.

The Storage Cell.—One of the most important chemical effects of the electric current is to be found in the storage cell. You of course know that this cell does not store up electricity, but simply converts electrical energy into chemical energy, which may be changed back again into electricity at our convenience.

To make a simple storage cell place in a tumbler a dilute solution of sulfuric acid, made by pouring 25 c. c. of the concentrated acid into 200 c. c. of water. Sand-paper two strips of heavy sheet lead and suspend them in the solution. Connect these with a source of current. If you have direct current, connect in series with the cell 2 lamps in parallel. If you depend on cells, use either 2 bichromate cells or 3 dry cells. Let the current pass for 5 minutes. Then disconnect the charging battery and try the action of the storage cell on an electric bell. Does it ring? You will find that the action soon ceases. Charge it again, and, as you do so, note that bubbles of gas rise around each of the lead strips. At the anode oxygen escapes and at the cathode hydrogen. Remove each strip. You will find the anode covered with a chocolate colored deposit

of lead peroxide. This is a new chemical compound which has been formed by the action of the electric current. The cathode you will observe remains unchanged. When the battery is disconnected and the cell is placed on closed circuit a reverse set of chemical changes occurs, which sends a current in the opposite direction to the charging current.

You may wish to make a *real storage cell*. If you do, obtain from a local electrical store a regulation container and two *lead grids*. Make a stiff paste by mixing the lead oxide known as *litharge* together with dilute sulfuric acid and glycerin. With a small wooden paddle press this paste into the perforations of the lead grids. Insulate these two grids from each other at top and bottom with pieces of heavy rubber tubing and bind the two together with strong rubber bands. Prepare a solution of sulfuric acid of the same strength as that used in the simple cell. Fill the container nearly full of this solution and set into it the two grids. To each grid attach a copper wire and connect the cell to a source of current. If the cell is to be well charged, the current must run for a number of hours. Of course it would not be profitable from a practical standpoint to charge storage cells from an ordinary battery. But if you make a number of these cells and connect them together in series, you will be able to charge them from your direct current lighting circuit, using the lampboard rheostat to get the proper amount of current. Or you may have them charged at a battery service station. I think, however, that you will

be able to meet your electrical requirements without home-made storage cells. But it will be a simple matter for you to charge your own and the radio batteries of your friends by means of your rheostat and source of direct current.

The Chemical Action of a Simple Cell.—The simplest cell that you can make consists of strips of copper and zinc placed in a dilute solution of sulfuric acid. When this cell is on open circuit, examine the copper strip and note that there are no bubbles on it. Probably there will be bubbles about the zinc for reasons which we shall presently explain. Now bring together the wires from the two plates and watch the copper plate again. Almost immediately you will observe a myriad of fine bubbles rising in a sheath about the plate. These are bubbles of hydrogen gas. On closed circuit, the zinc goes into solution driving the hydrogen of the acid over to the copper. This hydrogen is in the form of little positively charged particles called *ions*. These ions deliver their positive charges to the copper plate and then bubble off as hydrogen gas. As the zinc ions from the other plate go into solution they carry away with them positive charges, therefore leaving the plate itself with a larger quantity of negative electricity than positive. We thus see that the copper plate comes to be charged positively and the zinc plate negatively.

There are two other characteristics of the electric cell which must be mentioned. They are *local action* and *polarization*. Local action is the solution of the

zinc because of impurities in it. This occurs both on open circuit and closed circuit and wastes both the zinc and the acid. Each little particle of impurity, usually carbon, sets up with a nearby particle of zinc a tiny cell, and there are myriads of these cells. To prevent this action, you must cover the zinc with mercury or use chemically pure zinc. *Polarization* is the greatly diminished current resulting from the accumulation of discharged hydrogen bubbles on the positive plate. This decreases the voltage of the cell and increases its resistance, both of which diminish the current. The bichromate and dry cells provide oxidizing agents which partly dispose of the hydrogen and eliminate to some extent the effects of this action.

There is a host of chemical industries which depend upon the electric current for their success. With electric energy from Niagara Falls in huge chemical plants girding the brow of the great cataract are made carborundum, Acheson graphite, calcium carbide, aluminum, metallic sodium and potassium, calcium carbide, Alundum, caustic soda, chlorine, hydrogen, oxygen, calcium cyanimide, and many other products of immense importance. The vision of Sir Humphry Davy, who first applied the electric current to the solution of chemical problems, has come to a tremendous realization.

One of the most interesting applications of the electric current occurs in the refining of copper, silver, gold, and platinum. Considerable quantities of the silver, gold, and platinum which we use are electrically sepa-

rated as "impurities" in connection with the metallurgy of copper and lead. To go through the refining departments of one of these large smelters and see the separation of the precious metals in quantities that would make you richer than Croesus is a most interesting experience. But we cannot describe these processes here. We must hurry on to other subjects equally as important to the home-laboratory worker as any that have preceded.

CHAPTER 26

SIR WILLIAM PERKIN

HAD anyone wandered into the home laboratory of Sir William Perkin—then just plain Bill Perkin—during the Easter vacation of 1856 he would have observed this lad hard at work over a dirty black mess of coal tar. He was endeavoring to extract the drug quinine from this mixture of carbon compounds, a problem which had been assigned him by his professor of chemistry at the Royal College of Science in London. One evening just at the close of a disappointing day's work, young Perkin poured alcohol into a mixture of aniline oil and other chemicals with which he had been working. To his amazement and delight, there flashed into view a beautiful purple dyestuff. The color fascinated him. He wondered what should have caused it. He resolved immediately to separate this compound and to discover the method of its preparation. For days he worked upon the problem, and, although it baffled him for a time, he would not give up. At length he discovered the secret and the result was mauve, the first of the aniline colors.

Here was Perkin, still a freshman in college and only eighteen years old, making one of the greatest discoveries in the history of chemistry. So enthusiastic did

he become that he decided to leave college and begin the manufacture of the new dyestuff. His father assisted him financially, but where was the machinery for carrying out his process? It did not exist. Perkin, not yet out of his teens, was compelled to design it and supervise its manufacture. To add to his difficulties the aniline oil necessary for his purpose was not to be had. He must prepare it from benzene. And again strong nitric acid could not be bought on the market. He had to distil it from Chili saltpeter and sulfuric acid. But Perkin solved every problem and had soon founded a vast new industry. He discovered other coal tar dyes and his business grew until he became a wealthy man. And then what did Perkin do? He abandoned chemical manufacture and built himself a private laboratory where he might spend the rest of his years in the delightful work of original research. And working there quietly in his laboratory year after year, Perkin added vastly to our knowledge of organic chemistry. So distinguished did he become that his king knighted him and he was honored by the leading scientific societies of every land.

CHAPTER 27

THE CHEMISTRY OF LIGHT

WHILE I was in college I served for three years as laboratory assistant in chemistry. One day I was assisting my professor with some experiments involving the chemical effects of light. In a small flask we prepared a mixture of equal volumes of hydrogen and chlorine gases. We sealed this tightly and placed it under a very heavy battery jar, one whose sides were fully a quarter of an inch thick. In front of this we ignited a heap of flash powder. There was a sharp explosion. We repeated the experiment a number of times. It seemed to be perfectly safe and we enjoyed the fun. It was equal to a whole Fourth of July celebration. Then of a sudden there came an explosion more violent than the others. The big battery jar was blown to fragments and the flying glass cut us, but not seriously. We were almost dazed. We had gotten much more than we bargained for. We might have lost our sight or even our lives. We had staged altogether too striking a demonstration of the chemical effects of light. I have repeated the experiment once since, and with similar results, although I took great pains to protect myself by placing a heavy glass screen between me and the explosive mixture. I consider it too dangerous ever to repeat again.

Hydrogen and chlorine have an exceedingly strong affinity for each other. Under the influence of heat or the chemical rays of light they unite with explosive violence. The light from burning magnesium, the combustible substance in flash powder, is particularly rich in these chemical rays. Other chemical reactions are brought about through the energy of light. The trees and plants, breathing through their leaves the carbon dioxide of the air and absorbing water through their roots by means of sunlight, unite them into woody fiber, stalk, and flower. In the heat of burning coal we are but utilizing the products of the chemical energy of the sun's rays stored up for us in past geologic ages. The marvels of modern photography would be impossible without the chemical effects of the violet rays of the solar spectrum.

Blue-printing.—Blue-prints, so essential to modern architectural and engineering work, are produced by chemical reactions wrought through the agency of sunlight. If possible, obtain from some blue-printing firm a yard of blue-print paper. Cut out a small square, lay upon it a key or some other opaque object, and expose it to the sun's rays for about a minute. You will note that the green color of the paper begins to turn blue. Then place it in water and a deep blue will appear wherever the sunlight has fallen, but the part covered by the object will become white. The washing at once develops and fixes the print. The water completes the chemical changes started by the sunlight and at the same time dissolves out the unaffected salts

in the portion of the paper covered by the object. When the print is dry it will be permanent. It will not disappear under the influence of sunlight. Even such an accomplishment as this would have been hailed with joy by the pioneers in the quest for the secrets of photography.

With india ink make a drawing or sketch some design on the oiled paper used by draftsmen. Place this over a square of blue-print paper and repeat the above process.

To understand *the chemistry of blue-printing*, it will be necessary for you to do some preliminary experiments with iron salts. Prepare rather dilute solutions of ferric and ferrous salts. Chlorides, sulfates, or nitrates will be satisfactory. You will also need solutions of potassium ferricyanide indicator and the reducing agent oxalic acid.

Fill a test tube half full of water and add a few drops of the ferrous solution. Follow this with a few drops of potassium ferricyanide indicator. You will obtain a deep blue color. This is a test for a ferrous salt. Now repeat the experiment, but substitute the ferric solution for the ferrous. This time you will obtain an olive green, varying in shade according to the amount of ferric salt present.

Now fill a test tube one-sixth full of ferric solution. Add an equal quantity of oxalic acid (it should be a saturated solution), dilute the contents of the tube well with water, and add a few drops of potassium ferricyanide. Expose this tube to the direct rays of the

sun. Very soon you will observe streaks of blue appearing in the tube, and gradually the whole solution will change color. If the color is not a good blue but of a greenish cast, you have used too much of the ferric salt. You may have to experiment some to get the right proportions.

What has happened? The color of the solution was first green, owing to the presence of the ferric salt. Ferricyanide gives a blue only with a ferrous salt. Therefore, if a blue color appears, it must be that the ferric salt has been changed, or reduced, as we say, to ferrous form. And that is just exactly what has happened. The chemical rays of the sun have produced this change, but let us examine this further. We used oxalic acid. Did that have anything to do with the chemical reaction? To discover the answer, prepare another test tube exactly as before, but leave out the oxalic acid. Expose it to the sunlight. You will find that no color change occurs. Evidently the chemical rays of the light are powerless without the assistance of the acid. The oxalic acid acts as a medium to bring about the change. We have still to learn whether the color change might not be produced without the aid of the sunlight. Prepare a test tube exactly as you did the first one, using ferric salt, oxalic acid, and indicator, but this time place it in the dark. After a half hour examine the mixture. No change occurs. Therefore we see that the blue-print reaction consists in the reduction of a ferric salt to ferrous form through the agencies of a reducing agent and sunlight.

Preparation of Blue-print Paper.—If you wish to prepare blue-print paper, two solutions will be required. The substances and their proportions are given below.

Solution I:

Ferric ammonium citrate,	40 g.
Water,	200 c.c.

Add ammonia water until a decided color change occurs.

Solution II:

Potassium ferricyanide,	40 g.
Water,	200 c.c.

Prepare a mixture of equal parts of the two solutions and for every 100 c.c. of the mixture add 30 c.c. of a saturated solution of oxalic acid. The mixture must be kept in the dark.

Obtain a good quality of unglazed paper and draw squares of it across the surface of the mixture, distributing the liquid as evenly as possible, but keeping the upper side dry. Let these squares dry in the dark. Press them, if you like, between two smooth surfaces of wood upon which you have placed a heavy weight. Like all blue-print paper, they will have a greenish color before exposure to the light. The process of making a print is exactly similar to that with the regulation paper already described.

Silver Salts in Photography.—The problem of causing the rays of the sun to paint our pictures was a most baffling one. The Frenchman Daguerre solved it. He had exposed in his very crude camera of that

day a silver plate covered with a thin film of silver iodide, prepared by exposing the plate to the vapor of iodine. The chemical rays of the sun produced an imprint of the object focused upon it, but he could find no means of developing the image. He worked upon the problem for years without success. Then one evening, while he was engaged upon this fascinating quest, he was unexpectedly called from his laboratory. Quickly he thrust into a cupboard the silver plate upon which he was working. Imagine his delight, when the following morning he opened the cupboard and found a beautiful picture. During the night in some mysterious way the image had been developed. But how? That was a question which he must answer. His first step was to expose another plate in his camera and leave it in the cupboard the following night. Again the picture was developed. Then he began his search. He tested every chemical in the cupboard. At last he discovered that the plate had been placed near to a dish of mercury and that the vapor of this liquid metal had brought out the image. He quickly perfected the process and in a short time was making the first real photographs ever produced.

To learn the chemical effects of light on silver salts and the reactions which take place in the developing and fixing of a plate, film, or print, you will need four solutions:

36 g. of potassium bromide per liter of distilled water.

17 g. of silver nitrate per liter of water.

250 g. of "hypo" (sodium thiosulfate) per liter of water.

10 g. of hydrochinone, 20 g. of sodium sulfite, 1 g. of potassium bromide, 1 g. of citric acid, 20 g. of sodium carbonate to 1600 c.c. of water.

This last solution is the developer.

The silver nitrate and developer should be kept in dark colored bottles or away from the light. Since these volumes will be larger than you will require, you may prepare only one-tenth of the quantities indicated.

Fill a test tube one-fourth full of water and wrap it in heavy dark paper. Black is best. From burettes or a small measuring cylinder add 1 c. c. each of silver nitrate and potassium bromide solutions. This will precipitate silver bromide, one of the most sensitive salts to the effects of sunlight. Gently shake the tube to mix the precipitate with the liquid and removing the paper expose it to the action of the direct rays of the sun for 2 minutes. There may be some discoloration. This is due to the presence of organic water.

Pour into the test tube 5 c.c. of developer and continue to expose to the sunlight for 3 minutes. You will note that the contents of the tube rapidly turns dark. Add 10 c.c. of hypo and with your thumb over the mouth of the test tube invert it several times to insure mixing. The contents of the tube begins to clear and black particles appear. The sunlight and developer reduced some of the silver bromide to metallic form and the hypo has dissolved the unaffected silver salt. The black particles are metallic silver.

To prove this pour the liquid through a filter paper in your funnel and wash the black residue by passing repeated small portions of water through the filter. Place the portion of the filter paper containing the black residue in a test tube and add dilute nitric acid. The black particles will dissolve. Filter the solution into a clean test tube and wash it through with a little water. Add to the filtrate a few drops of hydrochloric acid and you will obtain a white precipitate, which will dissolve in ammonia water. This is the test for silver.

In a similar way prepare another test tube of silver bromide but this time keep it well wrapped in black paper. Add the developer as before and allow it to stand unexposed to the light for a few minutes. Then add 10 c.c. of hypo and remove the wrapping. You now see no black particles. The developer alone without the aid of sunlight is unable to reduce the silver bromide.

When a plate or film is exposed in the camera to the action of the light a chemical change is started which the developer is able to take up and continue. This action is greatest on those portions of the plate which have received the most light. Thus the light portions of the object, which reflect the most rays into the camera, will appear dark on the plate, while the dark portions will appear light. In this way we get a "negative." Yet, as you see, when the picture has been developed, there is still a considerable quantity of unchanged silver bromide. This must be removed or when taken into the light the picture will darken and

disappear. Therefore we immerse the developed plate in hypo which dissolves the unaffected silver bromide. After washing and drying, the plate is ready for use. As many of you know, the process of making a print is exactly the same as the one just described. But now we reverse the lights and shades and therefore obtain a positive.

Making a Photograph without a Camera.—This stunt is not as difficult as you might at first think. To carry it out you will need what is called “self-toning” paper. Such prepared paper already contains a developer and needs only “fixing” with hypo, washing, and drying. You can easily obtain this paper from any dealer in photographic supplies. Expose a small square of it to the sunlight. It will quickly turn to a deep brown except the portion protected from the light by your thumb or finger.

From some book select a picture which you wish to reproduce. The reverse side should be plain, and the paper should be thin enough so that, when you hold it to the light with the plain side toward you, the picture may be clearly seen. Under the picture place a piece of self-toning paper with the glossy side up. Cover the two with a glass plate and expose them to the sunlight for a few minutes. The time will vary from 5 to 10 minutes according to the thickness of the paper and the brightness of the light. Remove the paper and immerse it in a solution of hypo for 10 minutes. Then wash the print by placing it in a large basin of clean water for about 20 minutes and dry it.

To obtain a positive, place this negative face down on another piece of self-toning paper and expose to the sunlight again. Follow this by fixing, washing, and drying, and you will obtain a print that will surprise and delight you.

Action of Light on a Copper Compound.—Prepare a strong solution of cupric chloride. Immerse in it a sheet of polished copper. This treatment will precipitate on the copper a film of white cuprous chloride. Wash the plate in running water. Then place over it some design cut from heavy paper and expose to sunlight. The uncovered portion will be darkened by the sun's rays, while the protected portion will remain white.

Sunlight is a Powerful Antiseptic.—Direct sunlight will kill the germs of tuberculosis, diphtheria, and typhoid in from 6 to 8 hours.

CHAPTER 28

THOMAS A. EDISON

NEWSDOY and “candy butcher” on the Grand Trunk Railway at fourteen, most expert telegrapher in the Western Union System at twenty-two, world-renowned inventor before he was thirty, and one of our foremost scientists ever since—such has been the career of Thomas A. Edison, who as a lad spent every spare moment experimenting in a home-laboratory.

So interested was young Edison in performing experiments that he turned the baggage car of his train into a laboratory and print-shop. There he published a small paper giving the very latest news, which he obtained from telegraph operators along the way. At one time he sold as many as five hundred copies of each issue at three cents a copy. But Edison’s fondness for experiments more than once got him into trouble. One day while at work in his baggage-car laboratory a bottle of phosphorus was jarred to the floor and immediately blazed up setting the interior of the car on fire. In the midst of his trouble who should appear but the conductor, an irate Scotsman, who had but little sympathy with Edison’s laboratory “whims.” Seizing a pail of water he dashed out the fire and then at the

first station stopped the train and pitched Edison off, laboratory, print-shop and all.

About this time Edison snatched the three-year-old son of one of the Grand Trunk telegraph operators from in front of an approaching express train, narrowly escaping with his own life. To reward him the father taught Edison telegraphy. Although he was then but seventeen, Edison applied for and received the position as night operator in his home town. His father allowed him to have a laboratory on the third floor of the house and in it he persisted in working during the daytime instead of sleeping, as he should have done. As a result, Edison frequently fell asleep at night and was severely censured by the train dispatcher. After a while the dispatcher hit upon the idea of having Edison send over the line to him the signal of the letter "A" every half hour, thinking that this would surely keep him awake. It worked beautifully for a few nights. Then Edison saw that he must devise some scheme to get more sleep, for he would not give up his laboratory. One evening he appeared at the station with a queer looking contrivance, which he had constructed in his laboratory during the day. With some wire he connected it to the telegraph instruments on the desk and the clock on the wall. Then he sat down to see what would happen. When the half-hour came, right on the dot, the device closed the circuit and sent a very good imitation of the letter "A" over the wire. He waited another half-hour and, satisfied that the mechanism was perfect, went to sleep.

The dispatcher thought he had become a model operator, until one night when he happened to be at the next station down the line. Hearing the signal come through as usual on the half-hour, he pulled out the switch and called Edison, thinking he would have a chat with him. But to his amazement no one answered his repeated calls. After fifteen minutes he rushed out of the office and jumping upon a hand-car hastened to the rescue. Arrived at the station, he peered anxiously in at the window more than half expecting to find Edison dead of heart failure or killed by bandits. Imagine his surprise to find the "model operator" peacefully sleeping on a couch in a corner of the office. His first impulse was to give him a rude awakening, but spying the curious looking contraption on the desk, he sat down to see what would happen. The dispatcher did not have long to wait, and although he had to admire the ingenuity of the young inventor, he was compelled to dismiss him.

Edison then entered upon the life of a wandering telegraph operator, drifting from place to place, and gradually becoming the cleverest man in the service. At length he found himself in Boston where he remained for a short time and then went to New York, arriving penniless. He borrowed money to buy a breakfast, and then walked into the operating room of a central office for the distribution of stock quotations to the brokers of the financial district. He had not much more than entered when the main telegraph instrument which sent out the reports broke down.

Pandemonium set in. Three hundred boys, one from every broker's office on the street, came rushing in to know what the trouble was. The manager of the office lost his head, and the superintendent was frantic. Edison quietly walked over to the instrument, located the trouble, and volunteered to fix it. The superintendent excitedly directed him to do so and to be "quick." Deftly Edison removed a broken spring, which had fallen between the gear wheels, and set the machine in operation. The boys disappeared and silence reigned once more. But the superintendent called Edison into his office and made him foreman of the company at three hundred dollars a month. The suddenness of his change of fortune almost took Edison's breath away, but he resolved to make good. And he has been making good ever since.

Here Edison began his wonderful career as an inventor. He had soon patented a number of improvements on the stock tickers of that day. Then the president of the company, General Lefferts, called him in and asked how much they were worth. Edison says he had intended to ask five thousand dollars and come down to three thousand if necessary. But when the time came, he did not have the nerve to ask so much and so he says, "General, suppose you make me an offer." "How would forty thousand dollars strike you?", General Lefferts replied. Edison says this caused him to come as near fainting as he ever did, but he managed to say that he thought it was fair.

He received his money and opened shops in New

Jersey where he has invented systems of duplex and quadruplex telegraphy, the phonograph, the incandescent lamp, motors and dynamos, the moving picture machine, his famous storage cell, and much more.

The supreme ambition of Edison's life was to become an inventor, and it came from his early work in a home laboratory.

CHAPTER 29

SPECIAL TESTS

AND now we are coming to some real chemical analysis. That is what you have been waiting for. You want to learn how to analyze something—to know how the chemist is able to tell just what substances are present in some unknown sample. Nothing is more fascinating than this kind of work. But it calls for the utmost care, keen observation, and sound judgment. If your life were to depend upon the testimony of a chemist in a murder trial, you would want him to be pretty sure of his facts. Of if you were to be prosecuted for selling food containing adulterants or illegal preservatives, you would not want the chemist to be doing any guess-work in the matter.

The Two Common Alcohols.—We have heard a great deal about alcohol in the last few years and in particular about the frightful effects that result from carelessness or ignorance in the use of wood alcohol. We can find no better starting point in our work of commercial analysis than to learn how to distinguish chemically between wood alcohol and grain alcohol.

Let us begin with a consideration of grain alcohol, or ethyl alcohol, as it is known in chemistry. This alcohol results from the fermentation of sugar. It

has been known from the earliest times, and, although its use as a beverage has been quite rightfully prohibited, the industrial uses of alcohol are just as important as ever. It is used in medicines, drugs, and extracts. It is an excellent solvent for organic substances. It is used in the manufacture of paints and varnishes. It is a fuel. By its oxidation we obtain acetic acid, and when distilled with bleaching powder the product is chloroform. It is the starting point in the preparation of ether. And there are many more uses. Therefore, you see, that this alcohol is one with which you will often meet, and it becomes of importance to be able to detect its presence with perfect certainty.

To learn the test for grain alcohol pour a cubic centimeter of it into a test tube, denatured will do, and add an equal quantity of sodium hydroxide solution. Warm the mixture gently and add iodine, a crystal at a time until you obtain a yellow precipitate. Smell of the contents of the tube. The odor is familiar. Where have you detected it before? I think you will soon remember that it is the odor so common about a hospital. The substance is iodoform, one of the most powerful antiseptics known to surgery. This test is a delicate one. Very small quantities of alcohol will give it.

Repeat the test on witch hazel, bay rum, extracts, and medicines suspected of containing alcohol. If a liquid is highly colored, it is best to distil off a little of it and test the distillate.

Test for Formaldehyde.—Before we can make the test for wood alcohol, we must learn the test for formaldehyde. We do not test directly for wood alcohol but convert it first into formaldehyde. This test is similar to that of the doctor's treatment who could cure only fits. He first threw his patients into fits and then proceeded to cure them.

Into a glass of milk put one or two drops of a 5 per cent solution of formaldehyde. Into a test tube pour 2 c.c. of concentrated sulfuric acid and add a few drops of ferric chloride solution. Then, inclining the test tube somewhat, carefully pour down its side a little of the milk. The milk will float on the surface of the acid. Examine carefully the color at the dividing line between the two layers. You will observe a violet coloration. Even very minute quantities of formaldehyde in milk will give this test.

To learn the appearance when no formaldehyde is present, repeat the test with milk known to be free from this substance. This time you will find only a brown color instead of the violet.

Test for Wood Alcohol.—In making this test we must oxidize the alcohol into formaldehyde. To do this you will need a spiral of heavy copper wire. About one end of a lead pencil or small iron rod wind as closely together as possible turns of heavy copper wire. About No. 14 will be the proper size of wire. Make about a 2-inch length of these turns and leave the free end extending upward for 8 or 9 inches.

Fill a test tube a quarter full of wood alcohol and



Heating the copper spiral preparatory to making the test for wood alcohol.

set it inside a bottle half full of cold water (Fig. 31). In your Bunsen flame heat the copper spiral until it comes to a bright red heat. This forms copper oxide on its surface. Immediately thrust the hot spiral into the test tube of wood alcohol. Leave it until the boiling ceases. Then repeat the process six times. Note

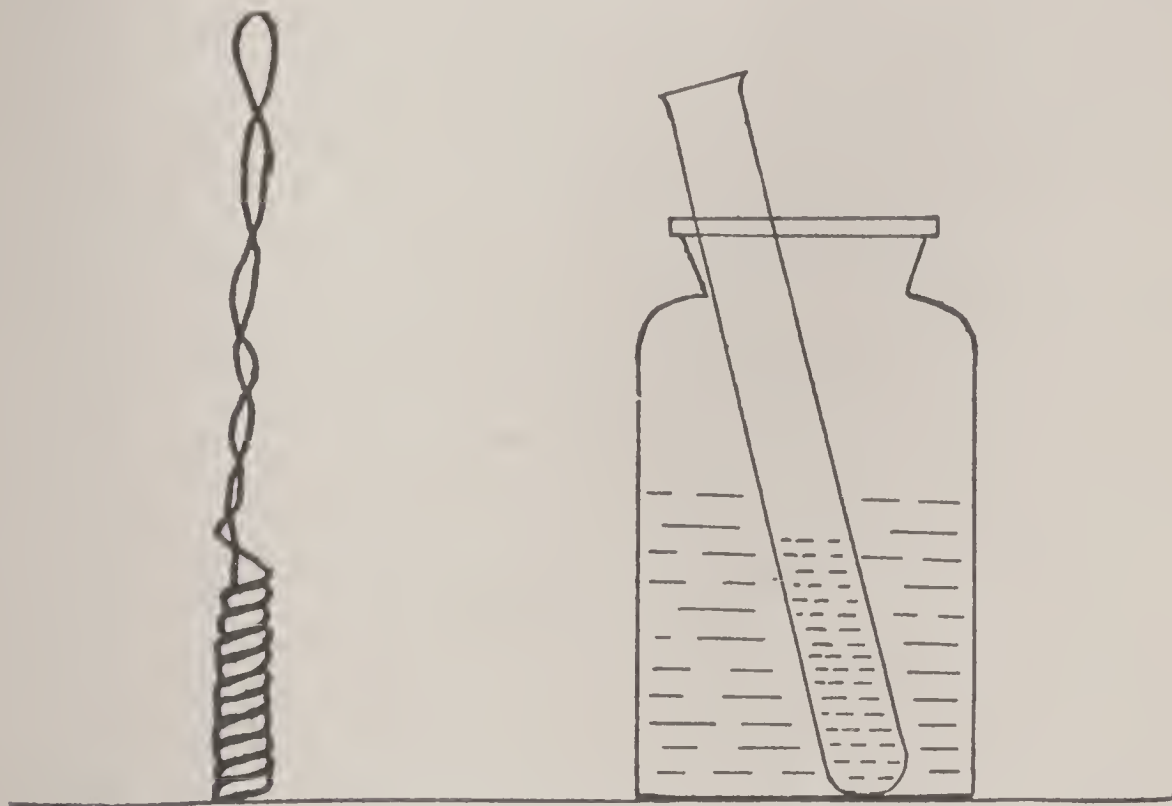


FIGURE 31.

Apparatus used in testing for wood alcohol.

that, as you remove the spiral, the copper has a very bright luster. The copper oxide which formed on the surface has been entirely reduced in oxidizing some of the wood alcohol to formaldehyde.

Into a test tube of milk pour a little of this oxidized solution of wood alcohol. Prepare another test tube containing 2 c.c. of concentrated sulfuric acid and a

few drops of ferric chloride solution. Then, as before, carefully pour some of the milk down the inclined side of the test tube onto the surface of the acid. Again you will obtain the violet color showing the presence of formaldehyde. This time the formaldehyde must have come from the wood alcohol.

This test may be made upon a beverage or any other liquid suspected of containing this deadly substance. I very often receive at my laboratory samples to be tested for wood alcohol. You can learn to make the test as well as an expert.

Food Tests.—There is no more important branch of analytical chemistry than that of food analysis. And this is a kind of analysis which the amateur can carry out very successfully. We often wish to know which of the three principal kinds of food are present, and sometimes whether adulterants have been used. Foods are classified as carbohydrates, proteids, and fats. *Carbohydrates* are those foods which contain starch and sugars, such as vegetables and cereals. “Carbo” means carbon and “hydrate” water. Therefore a carbohydrate is a compound containing the element carbon together with hydrogen and oxygen in proportion to form water. *Proteids* are food substances containing nitrogen. White of egg, casein of cheese, and lean meat are examples. The *fats* of course comprise butter, olive oil, salad dressings, and the fat of meats. We shall begin this work with the tests for these three principal food constituents.

Test for Carbohydrates.—We can make a general test for carbohydrates and separate tests for particular members of this class such as starch, glucose, and cane sugar. This test depends upon the fact that a carbohydrate will give a purple color with an alcohol solution of alpha-naphthol. This solution may be made by dissolving 3 g. of the alpha-naphthol crystals in 25 c. c. of alcohol. Your druggist will be able to supply this substance or obtain it for you.

Pulverize some cereal such as wheat, barley, or rice and boil it in a test tube two-thirds full of water. Filter this into another test tube and to the filtrate add 2 c.c. of the solution of alpha-naphthol. Then inclining the test tube pour slowly down its side 3 c. c. of concentrated sulfuric acid. The acid being heavy will slide to the bottom of the tube and between the layer of acid and that of the mixture above will appear a purple ring. This test may be repeated with potatoes, vegetables in general, sugars, and any other food which you wish to test for carbohydrates.

The Starch Test.—Boil a pinch of starch in a small beaker two-thirds full of water. Pour a little of this thin paste into a test tube and nearly fill it with water. Then prepare a dilute solution of potassium iodide and add to it a few crystals of iodine. One gram of iodide to 25 c.c. of water will be sufficient. Add a few drops of the iodide solution to the test tube containing the starch and you will obtain a deep blue color. This is the starch test.

Cut a potato and place a drop of the iodide solution

directly upon the freshly cut surface. A blue color will result. This test may be repeated with any other starchy food.

Starch in Ice Cream.—Starch is frequently used in ice cream. In small quantity this is desirable but in an excessive amount starch is an adulterant. Place a little of the ice cream in a test tube and boil it with 10 c. c. of water. Let it cool and add a few drops of the iodine solution. If starch is present, the tell-tale blue will appear.

Glucose.—Glucose is one of the two principal sugars and an important carbohydrate. It is used in great quantities in the manufacture of confections, and it is obtained commercially from starch for the manufacture of alcohol. To make the glucose test you will need two solutions known as Fehling's solutions. These solutions are prepared as follows:

Solution No. 1: 17 g. of copper sulfate in 250 c.c. of water.

Solution No. 2: 86 g. of Rochelle salts and 25 g. of sodium hydroxide in 250 c.c. of water.

Now dissolve a little glucose in a half test tube full of water and add 1 c.c. of each solution. A deep blue color will at first form. This color is characteristic of copper solutions in the presence of an alkali. But heat the contents of the tube and the blue will change to a bright brick red. This is the test for glucose.

You may easily test confections, honey, maple syrup, molasses, jam, and jelly for glucose. Dissolve a

little of the sample in water and apply the test as above.

Cane Sugar.—What we usually speak of as sugar comes from the cane. Try the glucose test on this sugar. You will be unable to obtain the red precipitate this time. Empty out the mixture, dissolve a little cane sugar in half a test tube of water, and add one drop of concentrated hydrochloric acid. Heat this a few moments at nearly the boiling temperature, moving the test tube back and forth through the flame so as not to develop a small geyser and shoot the contents of the tube onto yourself or the table. (This is always rutable in heating a test tube containing a solution.) Then apply the Fehling test. This time you will obtain the red precipitate of cuprous oxide showing that the cane sugar has been changed into glucose. The hydrochloride acid causes each molecule of cane sugar to unite chemically with a molecule of water and split into two molecules of the lower form of sugar. This process is known as the “inversion of cane sugar.”

Conversion of starch into glucose may also be accomplished in the same way. To 1 g. of starch in 10 c.c. of water add 3 or 4 drops of hydrochloric acid and boil for ten minutes. Again apply the Fehling solution and you will obtain the test for glucose.

Tests for Fats.—To test for a fat we must dissolve it in some liquid such as benzol or gasoline. To a few drops of olive oil in a test tube add 3 or 4 c.c. of gasoline and set the test tube in a beaker of hot water. Let it stand for several minutes, shaking it occasionally,

and then pour a few drops of the solution onto a folded filter paper placed in the funnel. After the gasoline has evaporated, examine the paper. A grease spot will appear, which can be plainly seen if held to the light.

Finely divide any food which you wish to test, warm it for some time with gasoline, and filter. (*Keep gasoline away from the flame, and do not warm the mixture over a flame.*)

Test for a Proteid.—To make this test you will need a special reagent. It is prepared by mixing copper sulfate and potassium hydroxide solutions. Dissolve 3 g. of copper sulfate in 100 c.c. of water and 10 g. of potassium hydroxide in another 100 c.c. Add the copper sulfate solution to that of the hydroxide a drop at a time until a faint blue color is obtained.

Place two porcelain dishes side by side on your table. In the first put a small quantity of finely divided peas or lean meat. Cover it with 5 c.c. of the above reagent, and in the second dish place an equal quantity of this reagent for purposes of comparison. The solution in the dish in which you are making the test will become pink or violet in color. This coloration will be readily seen by comparing the solutions in the two dishes. This test may be repeated on any food-stuff.

Baking Powders.—The purpose of baking powder is to liberate carbon dioxide gas and cause the bread or cake to become light and porous. All baking powders contain sodium bicarbonate, which is ordinary baking

soda, and some acid substance to react with it and liberate the carbon dioxide.

Add a little water to some baking powder in a test tube and pour the gas that forms downward into another test tube containing 2 or 3 c.c. of limewater. Shake the limewater tube and note the white precipitate. This is proof, as you know, of carbon dioxide.

Three acid compounds are used in the preparation of baking powders. They are cream of tartar, alum, and calcium acid phosphate. Cream of tartar is best, for it is a fruit acid, coming from grapes, and not harmful. Dissolve a little cream of tartar in water and test the solution with blue litmus paper. The red color shows the acid character of this salt. Add to the solution a pinch of baking soda and note the vigorous effervescence of carbon dioxide. Make a mixture of about equal parts by bulk of baking soda and cream of tartar. Place some of it in a test tube and add water. The result shows that you have real baking powder.

Make two other mixtures, one containing baking soda and finely pulverized alum and the other baking soda and acid phosphate, known as monocalcium phosphate. Add water to these, and again you will obtain carbon dioxide.

Testing for Tartrates.—To learn this test dissolve a crystal of silver nitrate in 5 c.c. of water and add 2 drops of ammonia water. Shake a little cream of tartar in half a test tube of cold water. Filter and add 1 c.c. of the filtrate to the silver nitrate solution.

You will obtain a beautiful mirror of silver upon the inside of the test tube. Repeat this test with baking powders. You can readily tell whether or not they contain cream of tartar. Shake about 4 g. of the baking powder with a test tube of water, filter, and make the silver nitrate test as described.

Testing for Alum.—Most of the cheap baking powders contain alum. Dissolve a little alum in water and test the solution with blue litmus paper. Do you not see why it can be used to liberate carbon dioxide?

To learn one test for alum scoop out a small hole on the end of a charcoal stick. Place in it a little alum and moisten with a drop of water. Adjust your Bunsen burner flame so that it is about 2 inches high and colorless. Holding the charcoal stick in the left hand at the side of the flame and a blowpipe in the right, direct the flame upon the alum and heat it vigorously for several minutes. Then moisten the residue with a few drops of cobalt nitrate solution and reheat with the blowpipe. After the second heating, you will obtain a blue color. Repeat this test with a baking powder. If alum is present, the blue color will appear.

Another test for alum may be made by heating a small quantity of baking powder as strongly as possible in a porcelain crucible for some time. Use the full heat of the burner. Place the residue in a beaker and add boiling water. Let the beaker stand for a few minutes and filter, catching the filtrate in another small beaker. Add to the filtrate a strong solution of ammonium chloride, ordinary sal ammoniac. If alum is

present, a white "flocculent," or floating, precipitate will appear. You will also obtain the odor of ammonia.

Testing for Calcium.—If we find calcium present in a baking powder, we may be quite sure that it comes from monocalcium phosphate, so often used as the acid constituent. Shake a little of some baking powder with a test tube of water, filter and to the filtrate add a few drops of ammonium oxalate solution. If a white precipitate forms which is insoluble in acetic acid, but soluble in hydrochloric acid, calcium is present and undoubtedly the phosphate too. If you like, you may make the ammonium molybdate test for phosphates, described under the chapter on water.

Starch is usually present in baking powder as a filler and to take up the moisture absorbed from the air so that a powder kept in a box not air-tight will not deteriorate so rapidly. Make the iodine test already described.

Sulfates are always present in alum baking powders and may be in others. To test for them dissolve some of the powder in cold water, filter, and add hydrochloric acid to the filtrate a little at a time as long as effervescence continues. Then add a few drops of barium chloride solution. A white precipitate shows the presence of sulfates.

Ammonium salts are sometimes present as an adulterant. To 5 c.c. of the filtrate from a solution of baking powder add an equal quantity of a fairly strong solution of sodium hydroxide and heat. Hold in the

escaping steam a piece of moist red litmus paper. If the paper turns blue, ammonium salts are present.

Milk and Butter.—There is no more important article of food than milk, and yet it is frequently adulterated. You may easily prove the presence of all three of the principal food constituents in milk. To test for proteid it will be necessary to curdle some milk. Prepare the curd by mixing 25 c.c. of sweet milk with 175 c.c. of water and adding very slowly and with constant stirring 40 c.c. of a one and one-half per cent solution of glacial acetic acid. (Get your druggist to prepare this solution for you.) After 20 minutes, filter into a clean beaker and save the filtrate. To the curd upon the filter paper apply the test for a proteid already described.

Test some of the filtrate for sugar with Fehling's solutions. The red precipitate at once shows the presence of this carbohydrate.

The production of butter from milk proves the presence of fats.

Artificial coloring matter is often added to milk to give it a rich creamy appearance. To test for *annatto* stir into a sample of the suspected milk baking soda until it will just turn red litmus paper a faint blue. Then immerse a filter paper in the milk and allow it to remain for 12 hours. If annatto is present, the paper will be stained a reddish-yellow. A coal tar dye is sometimes used. To detect it add 10 c. c. of hydrochloric acid to an equal volume of milk and mix thor-

oughly. If *azo orange*, the color usually employed, is present a pink coloration will appear.

Milk Preservatives.—No other form of illegal adulteration is more common than to use chemicals to prevent milk from souring. The test for *formaldehyde* has already been given in connection with wood alcohol.

Ordinary *baking soda* is often used as a preservative. In a porcelain crucible evaporate 5 c.c. of the milk to dryness on a water bath or over a very small flame. Then, heating the crucible with the full heat of the burner, convert the residue into an ash. When the crucible has cooled pour upon the ash a few drops of hydrochloric acid. If you obtain effervescence, sodium bicarbonate is present. Save the residue.

Borax and boric acid are very often used as milk preservatives. Dissolve the residue from the previous test in as little water as possible and dip into the solution a strip of turmeric test paper. Holding the paper with tongs high above the flame dry it with very gentle heat. If either borax or boric acid have been used the paper will turn to a cherry red color. Upon the addition of ammonia, the color will become olive green.

Oleomargarine.—The simplest test for this very common substitute for butter consists in heating some of the sample in an iron spoon. Pure butter will melt quietly with the production of a large quantity of foam. Oleomargarine will sputter, but produces scarcely any foam.

You may also make the following test: Heat 50 c.c. of milk in a beaker nearly to boiling and add 5 g. of

butter to be tested. Stir the butter with a match stick until the fat has all melted. Then set the beaker in a basin of ice water and stir the mixture until the fat becomes solid. If the sample is oleomargarine, it will collect as a solid mass on the stick. If it is butter, the fat will remain suspended in the milk.

Coal tar dyes are often used to color butter. To 50 c.c. of water add a few cubic centimeters of alcohol, a teaspoonful of butter, and a pinch of cream of tartar. Immerse in the mixture some threads of white silk or wool and heat to boiling. If coal tar colors are present, the fibers will be dyed.

Saccharin is often illegally used as a substitute for sugar. It is a coal tar product and 500 times as sweet as cane sugar. To detect it in canned goods or any other foodstuff, shake the sample with a little water, filter, and add chloroform to 2 or 3 tablespoonfuls of the filtrate. Shake the mixture thoroughly and with a medicine dropper remove some of the chloroform, which, since it does not mix with water, will settle to the bottom. Place the chloroform in your evaporating dish and heat it to dryness on the water bath. Taste of the residue. If it is sweet, saccharin is present. Sugar will not dissolve in chloroform, but saccharin will.

To test the quality of vinegar evaporate almost to dryness a half teacup of vinegar in a small agate basin. Smell of the residue. If you obtain a distinct odor of baked apples, the sample is cider vinegar. If you get the odor of grapes, it came from wine. A malt vinegar

will give the odor of malt. If there is no residue, the vinegar has been made artificially from acetic acid, water, and a little flavoring matter.

Lemon extract often contains no genuine lemon oil. You can determine this by adding 2 c.c. of the extract to 50 c.c. of water. If lemon oil is present, there will be a milky appearance to the mixture. By the amount of this milkiness you will be able to judge of the quantity of lemon oil. If the mixture remains clear, no lemon oil is present.

Vanilla extract is often artificial. To determine whether it has actually been made from the vanilla bean, dissolve 3 g. of lead acetate in 10 c.c. of water and add an equal volume of the extract. If the extract is genuine, a heavy precipitate will form and settle to the bottom, leaving a clear liquid above.

The purity of olive oil may be determined by shaking 5 c.c. of the oil with 5 c.c. of concentrated nitric acid. If the oil is pure a pale green color will be obtained within a few minutes. If a brown, red, or orange color is obtained, the oil has been adulterated. Hold the test tube in a beaker of hot water for 5 minutes. If the oil is pure, the color will change to orange yellow. Peanut oil gives with nitric acid a pale rose color, which changes to brownish yellow on heating. Cottonseed oil gives a yellowish brown, changing to reddish brown.

Copper salts are often used in canned goods to give them a green color. To detect them strongly heat 10 g. of the sample in a porcelain crucible until an ash has

been obtained. Dissolve the residue by heating it with a little nitric acid. Evaporate the solution nearly to dryness, and dissolve it in 5 c.c. of water. Pour the solution into a test tube and add ammonia water. If copper is present, a deep blue color will be obtained.

Sulfites in canned goods serve two purposes. They bleach them and preserve them. Thoroughly pulverize 25 g. of the sample and put it into an Erlenmeyer flask. Add 5 g. of pure zinc and 10 c.c. of pure hydrochloric acid. Heat the flask gently and hold in the neck of it a strip of filter paper which has been moistened with a solution of lead nitrate or acetate. If sulfites have been used, the paper will become brownish to black in color.

Potassium nitrate, or saltpeter, is a common meat preservative. Get your druggist to prepare for you a small volume of a 1 per cent solution of diphenylamine in strong sulfuric acid. Put some of the meat in an evaporating dish and add to it 2 c.c. of this solution. If nitrates are present, a blue color appears immediately.

CHAPTER 30

GEORGE WESTINGHOUSE

GEORGE WESTINGHOUSE was another lad whose chief delight was in using tools. In the loft of his father's machine-shop at Schenectady, New York, he fitted up an "inventor's" laboratory in which, while still a boy, he built a rotary steam engine, with which he ran a boat on the Erie Canal. After he had become a man he patented this engine. But his father, a somewhat stern man of Dutch ancestry, did not approve of these boyish pursuits and wished George to work at the bench on worth-while projects. Still experiment and invent he would. One day his father led him to a pile of pipe, which he wished cut into pieces of a certain length. Wishing to escape the drudgery, George set his inventive brain to work and soon had a combination of tools which, when hitched to a power-shaft, automatically fed the pipe and cut it into the desired lengths. And all through life he could devise machinery to do exactly what he wanted it to do.

In 1861, when only fifteen, he attempted to run away and go to war. Foiled by his father, he waited two years and then enlisted. After the war he entered college, but study did not appeal to him, and he soon abandoned school for a life of invention. A

couple of railroad accidents led to his two most important early inventions. One was a patent car-replacer for quickly putting derailed cars back onto the track. The other was his world-famous air brake. The latter, one of the most important inventions ever made, was perfected before he was twenty-two. It brought him great wealth and made his name known on two continents. Thus at the very outset of his career he had supplied himself with abundant funds for further experimentation, and that was the passion of his life.

Under his direction the engineers in his immense laboratories and shops developed some of the epoch-making inventions of modern times. It was Westinghouse who saw, as no other man did, the future necessity for the manufacture and distribution of alternating-current power. His engineers brought out the transformer, the rotary converter, and the induction motor, indispensable to the electrical age in which we live. He electrified the World's Fair in Chicago in 1893 and built and installed the huge generators at Niagara Falls, the world's first big hydro-electric plant. He developed the interlocking block system of railroad signaling in America. On his private estate near Pittsburgh he put down the first of the big Pennsylvania gas wells and organized the Pennsylvania Gas Company.

From the days of his boyish contrivances to the end of his years Westinghouse was always an inventor, a wonderful man, equaled by few and surpassed by none in his chosen field.

CHAPTER 31

MORE REAL ANALYSIS

IN this chapter I wish to give sufficient instruction to enable you to do a great deal of original analytical work. When you have made the tests which I shall describe, you should be able to take any of the ordinary metallic compounds and determine with perfect certainty what metal it contains and from what acid it has been derived.

Some of these tests we have already made, but it will be best to summarize them with the new ones now to be given. When you identify a salt it is necessary to do two things—determine the metal it contains and the acid radical with which it is combined. We shall state first as briefly as possible the tests for the more important acid radicals. In every case you should repeat the test with a known sample of the salt.

Chlorides.—Dissolve a little sodium chloride in water and add a solution of silver nitrate. You will obtain a white precipitate of silver chloride which will not dissolve in dilute nitric acid.

Bromides.—For the bromide test you will need a solution of chlorine water. This must be freshly prepared. Before starting remember that chlorine is poisonous. Put your apparatus near an open window

or in a good draft and *do not breathe the gas*. But, in preparing it on the small scale which we shall employ, just be careful and there will be no danger. Fit a test tube to a one-hole stopper and delivery tube.

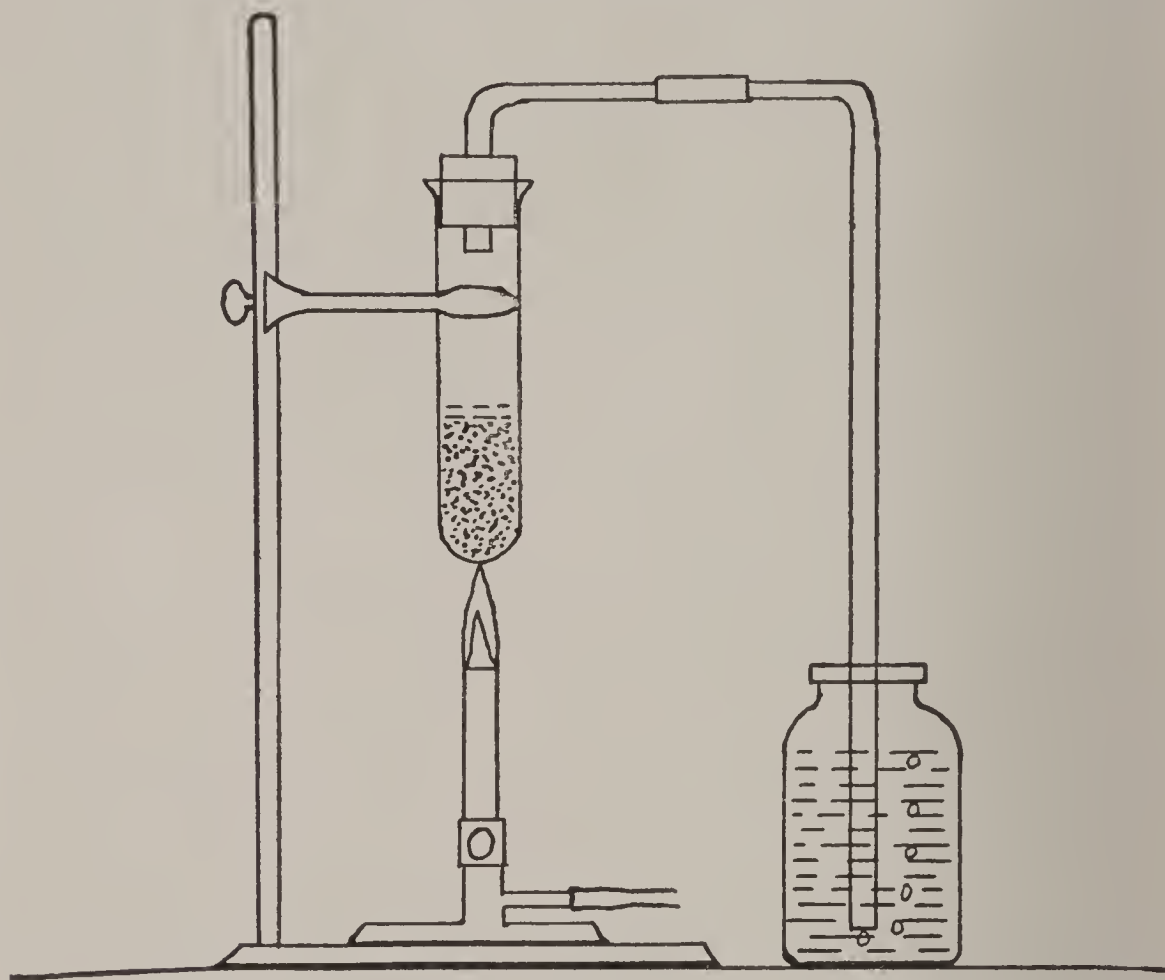


FIGURE 32.

Set-up for preparing chlorine water.

Clamp it to the ring-stand and removing the stopper place in the test tube 5 c.c. of concentrated hydrochloric acid. Sift into the acid a teaspoonful of manganese dioxide and replace the stopper. Let the delivery tube pass to the bottom of a small bottle filled with water (Fig. 32). Now heat the test tube very



Preparing chlorine water for use in testing iodides and bromides.

gently and a stream of chlorine gas will be driven off. Be careful not to heat the tube strongly enough to drive its contents over into the delivery tube. As the chlorine bubbles through the water, some of it dissolves and soon the water becomes greenish in color. When this happens, turn off the burner, remove the delivery tube and preserve the chlorine water in a stoppered bottle. The bottle should be dark colored, because otherwise the chemical rays of the sun cause the chlorine to unite with the water, and the strength of the solution becomes very weak.

To make the *bromide test* dissolve a little sodium or potassium bromide in half a test tube of water and add a few drops of chlorine water. Immediately you will observe that the solution turns brown. This is because the chlorine has set bromine free, since chlorine is more active than bromine. Now add a cubic centimeter of carbon disulfide and placing your thumb over the mouth of the tube shake the mixture vigorously. You will observe that the carbon disulfide becomes reddish in color and settles to the bottom of the tube. The bromine has dissolved in it and colored it. Any other bromide will give a similar test.

Iodides.—The test for an iodide is exactly the same as that for a bromide. Dissolve a particle of potassium iodide not larger than a pin head in two-thirds of a test tube of water. Add to the solution a few drops of chlorine water. Again a brown color appears. Add a cubic centimeter or two of carbon disulfide and shake. This time you will find that the carbon

disulfide becomes colored a rose to violet hue. If the color is very dark too much iodide has been used.

Instead of using carbon disulfide, you may also pour some of the solution to which you have added chlorine water into water containing starch paste. A deep blue color will appear.

Sulfides.—To a fragment of iron sulfide in a test tube add either dilute hydrochloric or sulfuric acids. You will very soon obtain the foul smelling odor of hydrogen sulfide gas. It is the odor with which you are already familiar as coming from bad eggs. But you must make a chemical test too. Hold in the escaping gas a piece of filter paper moistened with lead nitrate solution. Immediately the paper will become black. Lead sulfide has formed.

Sulfites.—To a half teaspoonful of sodium sulfite in a test tube add hydrochloric acid. You will get the sharp odor of sulfur dioxide. Pour the gas carefully down into a test tube containing a dilute solution of potassium permanganate. The purple color of the permanganate will immediately disappear.

Sulfates.—Dissolve a little of any sulfate in water and add a few drops of a dilute solution of barium chloride. A white precipitate of barium sulfate will form, which is insoluble in dilute hydrochloric acid.

Nitrates.—Prepare a rather strong solution of ferrous sulfate. Place 2 c.c. of it in a test tube and carefully pour down the side of the tube, held in an inclined position, an equal volume of concentrated sulfuric acid. Being heavy the acid will pass to the bottom and lift

the sulfate solution above it. Now pour down the side of the test tube a cubic centimeter of a dilute solution of sodium or potassium nitrate. A brown ring will appear at the dividing line between the two layers, which will widen by careful shaking and then disappear.

Carbonates.—To a little baking soda, washing soda, or marble in a test tube add hydrochloric acid and pour the gas into limewater. (Do not let the liquid pass into the limewater too.) Shake the tube and, as usual, you will obtain a white precipitate.

Phosphates.—To a solution of sodium phosphate add 2 c.c. of the ammonium molybdate solution, prepared in the work on water, and warm gently. A lemon-yellow precipitate will appear.

Oxalates.—Dissolve some sodium or potassium oxalate in water and add a solution of calcium chloride. A white precipitate will form which does not dissolve in acetic acid but will dissolve in hydrochloric acid. If an oxalate in solid form is heated in a test tube carbon monoxide gas will escape and burn with a blue flame at the mouth of the test tube.

Tartrates.—Heat some cream of tartar in a porcelain crucible until it darkens and gives the odor of burnt sugar. Dissolve another portion in half a test tube of water. To 5 c.c. of dilute silver nitrate solution add 1 c.c. of the cream of tartar solution and 2 drops of ammonia and warm. Metallic silver will be precipitated on the inside of the test tube.

Chlorates.—Heat a little potassium chlorate in a

test tube and thrust into the tube a glowing splint. It will be kindled into a flame.

Acetates.—Place enough sodium acetate in a test tube to fill the curved portion. Add 2 or 3 drops of water and an equal quantity of concentrated sulfuric acid. Warm and add a few drops of grain alcohol. You will obtain a rather pleasant fruit odor.

Permanganates.—Dissolve a single crystal of potassium permanganate in half a test tube of water and add a solution of sodium sulfite. The purple color of potassium permanganate will disappear.

Chromates.—To a solution of potassium chromate add barium chloride solution. A yellow precipitate forms which is soluble in hydrochloric acid. If alcohol is added to the solution, the color becomes green. Lead nitrate solution will also give chrome yellow with a chromate.

Borates.—To half a test tube of borax add 3 c.c. of concentrated hydrochloric acid and dip into the solution a piece of turmeric paper. Holding the paper high over the flame dry it. A cherry red color will appear which changes to olive green when ammonia is added.

Moisten some borax powder with hydrochloric acid in an evaporating dish and add alcohol. Ignite the alcohol, and in a few moments you will observe a green colored flame.

Tests for the Metals.—The tests just described include most of the acid radicals which you will encounter. We must now learn tests for the more

common metals. The work being described in this chapter is not systematic qualitative analysis, but it will give you an excellent start toward it.

Flame Tests.—Sodium, potassium, lithium, calcium, barium, and strontium may be detected by the colors that they give to the Bunsen flame. For this work you will need a platinum wire, although a nickel wire may be substituted. Platinum of course is expensive, but you will need a piece only 3 inches long. When you get it seal it into the end of a small glass tube about as long as the wire. Hold the end of the tube in the flame and thrust about a quarter of an inch of one end of the wire inside the tube. Slowly turn the tube and as the glass melts it will run together and make a perfect seal.

Dip the wire into concentrated hydrochloric acid and hold it in the outer, oxidizing, flame of the Bunsen burner. Repeat this once or twice to clean the wire. When the wire is clean, the flame should be colorless.

Have at hand solutions of the salts of the six metals named above. Dip the clean platinum wire into the solution of sodium and hold it in the outer flame. A deep yellow color results. This is the characteristic test for *sodium*.

Clean the wire as described above and dip it into a solution of *potassium* salt. If the solution is pure, you will obtain a violet colored flame. Often sodium is present, and, if so, even in minute quantity, the violet of potassium will be completely masked by the yellow of sodium. Therefore, whenever testing for

potassium, look at the flame through two thicknesses of blue glass. The glass will allow the violet color of potassium to pass through it, but not the yellow of sodium. Repeat the test until you are sure of it.

Make the test for each of the other four metals, cleaning the wire after each test. *Lithium* will give a beautiful carmine red, *calcium* gives an orange, *strontium* a bright red, and *barium* green.

Hold a piece of glass tubing in the flame until it softens. The yellow color of sodium is obtained, for sodium compounds are used in the manufacture of glass.

Borax Bead Tests.—A number of other metals are detected by the colors which they give when fused with a borax bead. For these tests, too, you will need the platinum wire. Make a small loop on the end of the wire, heat the loop in the outer flame of the burner and dip it into borax powder. Some of the powder will cling to the wire. Thrust this into the flame and heat it until the powder fuses and runs into a clear glassy bead which completely fills the loop. To obtain such a bead the loop must be small.

Touch the hot bead as lightly as possible to a crystal of *cobalt* nitrate, and reheat in the flame until the cobalt compound completely fuses with the bead. Allow the bead to cool and note that you have a deep blue color. Cobalt will always give this color.

To clean the wire, heat it in the flame and thrust the hot bead into water. With your fingers remove

the material on the wire and, making another loop, prepare a fresh bead. If the bead is not colorless, another must be made.

Using the substances in solid form, repeat the test with compounds of *copper, chromium, iron, nickel, bismuth, and manganese*. In each case characteristic colors will be obtained which you should note and record. In some instances the color is different when the bead is hot from what it is when cold. Manganese gives an amethyst red, but only a tiny speck of the compound should be used.

Cobalt Nitrate Tests.—In the end of a charcoal block scoop out a small hole and place in it a little of some *zinc* compound. Moisten this powder with a drop or two of water and, using a small Bunsen flame, heat it strongly with the blowpipe for several minutes. Moisten the residue with a drop of cobalt nitrate solution and reheat with the blowpipe. Zinc gives a green color after the second heating.

Repeat the test using compounds of *aluminum* and *magnesium*. Alum may be used for the first and Epsom salt for the second. In one case you will obtain a blue color after the second heating and in the other a rose color.

Other Blowpipe Tests.—Make a mixture of powdered blue vitriol and washing soda or baking soda. Heat some of the mixture on a charcoal block with the blowpipe. Globules of metallic copper will be obtained. *Silver* nitrate will give white shiny globules of silver.

Compounds of *lead* and *tin*, when heated in the same way, will give malleable beads of metal and a deposit on the charcoal. *Antimony* and *bismuth* give brittle beads and a deposit.

Iron Tests.—Prepare dilute solutions of any ferrous and ferric salts that you may have. For indicators prepare dilute solutions of potassium ferrocyanide, potassium ferricyanide, and potassium sulfocyanate.

Add a cubic centimeter of the ferric solution to two-thirds of a test tube of water and test with a few drops of potassium ferrocyanide indicator. You will obtain a deep blue color. The ferrous salt gives a light blue with this indicator.

Test both salts with each of the other indicators. Use very dilute solutions each time, for these are delicate tests. I think you will find that the ferrocyanide and sulfocyanate tests are best for a ferric salt and the ferricyanide test for a ferrous salt.

Many soils and rocks contain iron compounds. Wherever you see a red or yellow color, iron is probably present. To test for it place a little of the rock or soil in a test tube and boil it with concentrated hydrochloric acid. Filter and test the diluted filtrate with potassium ferrocyanide or potassium sulfocyanate. The color obtained will be the evidence.

Lead, Silver, and Mercury.—These metals form the first group in the systematic separation of the metals and their detection in qualitative analysis.

To 5 c.c. of a solution of lead nitrate add hydro-



Making a blowpipe test for a metal.

chloric acid as long as a precipitate seems to form. Allow the precipitate to settle and add another drop of acid to the clear liquid above. If all of the lead has been precipitated, no more precipitate will form. When this is the case, pour off the clear liquid. Fill the test tube with cold water, shake and let the precipitate settle. Pour off the clear liquid and fill the test tube with hot water. You will find that the precipitate, which is lead chloride, dissolves in hot water but not in cold.

Repeat this procedure with silver nitrate solution. When you have determined that the silver chloride, which forms, does not dissolve in hot water, add to some of the precipitate ammonia water. This readily dissolves it.

Make the same tests with a solution of *mercurous* nitrate. I say "mercurous" because there are two series of mercury salts and only the "ous" salt belongs in this group. I think you will find that neither hot water nor ammonia dissolves mercurous chloride. But you do find that ammonia turns it black.

Now you have the means of detecting and separating these three metals.

To separate lead, silver, and mercury place in a small beaker 5 c.c. each of solutions of lead, silver, and mercurous nitrates. Add a liberal quantity of dilute hydrochloric acid and filter into a clean beaker. To determine whether enough acid has been added put a few drops of it into the filtrate. If no more precipitate forms, enough has been added. But, if a precipi-

tate does form, add more acid and filter again. Repeat this until the precipitation is complete. Wash the precipitate upon the filter by passing cold water through it a number of times. Then pass hot water through and catch the filtrate in a clean test tube. From the previous tests, which chloride should be dissolved? Add to the filtrate a little potassium chromate solution. The yellow precipitate of "chrome yellow" proves that lead has been separated. There now remain on the filter the chlorides of silver and mercury.

To separate the silver and mercurous chlorides pour through the filter 5 c.c. of ammonia water catching the filtrate. A black residue appears on the filter. You will recognize this as the mercury compound. Add to the filtrate dilute nitric acid and a white precipitate will form. By the neutralization of the ammonia the silver chloride, which dissolved and passed into the filtrate, has been re-precipitated. You have now completed the separation of the metals of Group I.

Using the various tests given in this chapter you will find it most interesting work to identify unknown compounds. If you have a chum who is working with you, each may give to the other unknown substances to analyze and report upon.

Organic matter may be detected by heating the substance in a porcelain crucible. If such matter is present, the substance will char.

CHAPTER 32

CRYSTALS

CRYSTAL forms of matter are among the most beautiful to be found in nature. Diamond is crystallized carbon. Sir Humphry Davy burned a diamond in pure oxygen and obtained nothing but carbon dioxide. Others have repeated the experiment many times since. Snowflakes are crystals formed by the condensation of water vapor at temperatures below freezing. The mineral wealth of the earth has been deposited in crystal form during the countless ages of the past. Beautiful specimens of gypsum, quartz, Iceland spar, and alumina bear evidence of nature's handiwork. By what processes precious metals have come to be embedded in veins of crystal rock we can only guess. But crystals disclosing, under the microscope and often to the unaided eye, definite geometric shapes abound in the earth's crust and may be prepared by artificial means.

In the laboratory we may prepare crystals by cooling hot saturated solutions. The process of crystallization is one of the most important in the purification of chemical compounds. It was as the result of a prodigious amount of crystallization and recrystallization many times repeated that Madame Curie

separated from several tons of pitchblende a few hundredths of a gram of the element radium, a million and a half times more active than the parent mineral itself. The Chili saltpeter of South America and the potash from the famous Stassfurt deposits of Germany are purified by crystallization. Cream of tartar is separated by crystallization from the crude salt which deposits on the inside of grape juice casks. There are many other illustrations.

Preparation of Crystals.—Obtain some dry powdered washing soda. Heat 50 c.c. of water to boiling in a small beaker. Sift the washing soda into it with constant stirring as long as the salt will dissolve. When a little remains that will not dissolve after further boiling, allow the solution to cool. As it does so, crystals will form. After the solution has cooled for several hours, pour off the liquid, and place the crystals upon a filter or blotting paper.

When dry, heat a few of the crystals in a dry test tube, inclining the mouth of the tube downward somewhat. Note that water forms and that the salt loses its crystal appearance. Evidently the crystal form is due to the presence of water. This water is chemically combined with the salt and known as water of crystallization. Expose some of the crystals to the air for a few hours and observe that they become dull and white on the surface, losing their bright crystalline appearance. After a time you will be able to crumble them to a powder with your fingers. Such salts which give off their water of crystallization and

crumble to a powder upon exposure to the air are called *efflorescent salts*.

Dissolve as much *potassium nitrate* as possible in 25 c.c. of hot water. Allow the mixture to cool and you will obtain a mass of crystals. These two experiments also illustrate the fact that most salts are more soluble in hot water than they are in cold. Repeat with Epsom salt.

In a dry test tube heat a crystal of *copper sulfate*. Water will be driven off, the blue color will disappear, and, when cold, you will be able to crush the salt to a dry powder between your fingers. Dissolve this powder in a few drops of hot water, place it in your evaporating dish, and let it stand. The blue crystals will reappear.

Just as in these cases, many salts depend upon chemically combined water for their crystal forms. But there are others which do not. Heat some crystals of potassium chlorate in a test tube. The crystals melt, but no water is formed. Test all the salts which you have to see whether they will give off water and crumble to a powder.

If you have a balance, place in one pan a porcelain crucible and weigh it very carefully. Then weigh into it exactly 2 g. of some salt containing water of crystallization. Crystallized barium chloride is an excellent salt for this purpose. Place the crucible on a clay triangle over the Bunsen burner and heat first with a small flame, gradually increasing it to its full capacity. Continue the heating for 20 minutes. When

the crucible has thoroughly cooled, reweigh it. The weight is less now. Find the loss of weight and determine the percentage of water of crystallization that has been driven off. Repeat this experiment a number of times. You will find that the percentage loss is always the same. This fact proves that the water is combined in definite, or chemical, proportions. This *law of definite proportions* is one of the fundamental principles of experimental chemistry.

Sulfur Crystals.—Dissolve a piece of roll sulfur about as big as a pea in a little carbon disulfide. (*Keep the carbon disulfide away from a flame.*) Pour the solution onto a watch glass and let the liquid evaporate. Crystals with diamond shaped faces known as rhombic crystals will be left. If you examine them with a small lens their shape will be clearly seen.

Sulfur also forms crystals of another shape. Melt some roll sulfur in a test tube. Heat the tube very slowly and move it about in the flame. Remove it from the flame whenever the contents begin to look red or dark. It will gradually melt to a clear straw colored liquid. Pour this into a folded filter paper. Watch the sulfur as it cools and when the surface has nearly frozen over, open up the filter, pour out the liquid remaining in the center, and you will have a mass of needle shaped crystals. They are called monoclinic, or prismatic crystals. Examine them with a hand lens and you will see that they resemble long slender prisms in shape.

A Lead Tree.—Prepare a fairly strong solution of lead nitrate (about 10 g. to 100 c.c. of water) and place it in a cylinder or bottle. Suspend in the solution a zinc rod covered with a layer of asbestos paper. After standing several hours, you will find crystals of lead on the outside of the rod and running off into the solution like the branches of a tree.

Preparing "Gold."—Mix equal parts of powdered tin and flowers of sulfur with one-eighth their bulk of ammonium chloride powder. Place the mixture in a porcelain crucible and cover it with a thin layer of the ammonium chloride. Cover the crucible with its lid and heat it over the Bunsen burner. A white smoke will issue from the crucible due to the vaporization of the ammonium chloride and a mass of bright crystals resembling gold will appear on the crucible lid.

Preparation of Alum.—In each of two beakers place 200 c.c. of water. In one dissolve 35 g. of potassium sulfate and in the other 68 g. of non-crystalline aluminum sulfate. To dissolve them quickly heat the beakers gently and with constant stirring. When the salts have dissolved, pour the solutions into an agate basin, and evaporate off half to two-thirds of the water. Then let the mixture cool. A mass of crystals will form which consist of a chemical combination of the two salts. This double salt is true alum.

You may grow a large crystal of alum by suspending a tiny crystal by means of a thread in a saturated solution of the salt. Let the process continue for several days.

Supersaturation.—Melt 50 g. of photographer's hypo in a small Erlenmeyer flask or beaker. Allow the solution to cool without disturbance. Then add a tiny crystal of the same salt and at once the whole contents of the flask will become a mass of crystals. The salt melts and dissolves in its own water of crystallization. If undisturbed, the salt will remain in solution, even when cold. But upon the addition of a tiny seed crystal it takes back its water of crystallization and solidifies.

Crystals of Iodine.—Place in the bottom of a dry test tube 1 g. of potassium iodide mixed with one-fourth its bulk of manganese dioxide. Clamp the test tube in a vertical position to the ring-stand and pour carefully down upon the mixture without wetting the sides of the tube 2 c.c. of concentrated sulfuric acid. Warm the tube very gently. The violet vapor of iodine will at once appear and much of it will condense on the cold upper walls of the tube giving crystals of the solid. These may be preserved in a tightly stoppered bottle or test tube.

Chemically Pure Sodium Chloride.—As you know, sodium chloride is ordinary table salt. As we buy it at a grocery store, salt is not pure. You have noticed many times how salt absorbs moisture in damp weather. This is due to the presence of small amounts of calcium and magnesium chlorides. These are what we call deliquescent salts. If you have it, place a little of one of them on a watch glass and leave it for a few hours. The salt will become moistened on the surface,

and, if left for a day or two it will absorb enough water to form a solution. These and other impurities are often present in the commercial article. The readiness with which salt dampens is an indication, not only of the amount of impurity present, but also of the state of humidity of the air. It will be interesting to prepare crystals of chemically pure sodium chloride and compare its properties with those of ordinary salt.

Dissolve about 75 g. of ordinary salt in 180 c.c. of water, distilled if possible. Filter into a clean beaker to remove any dirt and obtain a clear solution. Add concentrated hydrochloric acid until a precipitate just begins to form. This is sodium chloride, which is being thrown out of solution. Now set up a hydrochloric acid generator just as was described in the chapter on acids, bases, and salts. Place in the generator a small handful of table salt, or better rock salt. Connect the delivery tube to a funnel clamped mouth downward to a support and dipping beneath the solution of sodium chloride held in a beaker or small battery jar (Fig. 33). Through the thistle tube of the generator pour concentrated sulfuric acid and apply heat *at once* using a small flame. The hydrochloric acid gas which passes into the solution being more soluble than the sodium chloride will drive this compound out of solution as chemically pure salt. Any impurities present in it will remain in solution.

When the precipitation seems to be complete, *first remove the funnel* and then turn off the gas. After the salt has settled, pour off the clear liquid. Cover the

salt with pure dilute hydrochloric acid and shake well. Allow the salt to settle again and pour off the liquid.

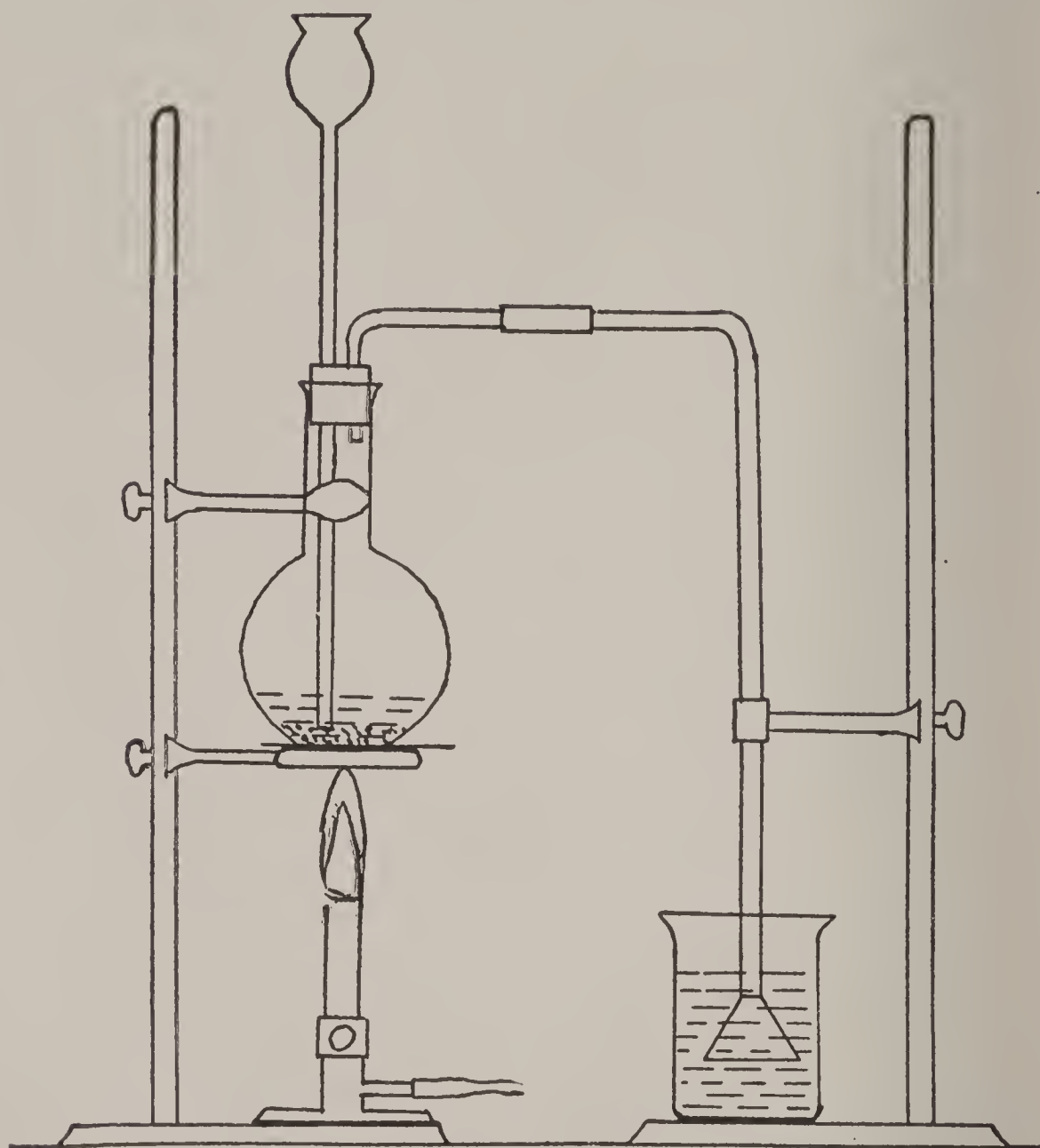


FIGURE 33.

The preparation of chemically pure sodium chloride by passing hydrochloric acid gas into a saturated solution of the salt.

Repeat this operation once or twice. Its purpose is to wash the crystals free from the liquid clinging to them.

Finally filter the salt and allow the crystals to dry. You now have chemically pure sodium chloride.

Place some of it on a watch glass and observe whether it becomes moist as ordinary salt does on exposure to the air. Examine some of the crystals with a lens and note their form.

A Crystal Garden.—Prepare a solution of ordinary water glass such as is used for preserving eggs. This solution is obtained by mixing the liquid with water until its specific gravity is 1.10. Specific gravity may be determined with a hydrometer. Possibly you can borrow one from the school laboratory or from a druggist. If not, mix the silicate with water until the solution is just a little heavier than water. If a little of it is poured into a test tube of water it should just nicely sink to the bottom. If you can weigh the solution, 100 c.c. of it should weigh 110 g. plus, of course, the weight of the container.

Place a pint of this in a battery jar or a large beaker and drop into it at various points on the bottom crystals of copper sulfate, cobalt nitrate, nickel sulfate, manganese sulfate, and ferrous sulfate. Leave the jar undisturbed for several days. Gradually there will appear from each substance a crystal growth of stalk and branches, each "tree" differing in color from the others.

Boiling by Freezing.—In a small flask melt 50 g. of *crystallized* sodium sulfate, known commercially as Glauber's salt. The salt will melt and dissolve in its water of crystallization. When the salt has entirely

melted, plug the neck of the flask with cotton wool and allow the solution to cool undisturbed.

Meanwhile make an ether thermometer. To do so blow a bulb on the end of a glass tube. You will find

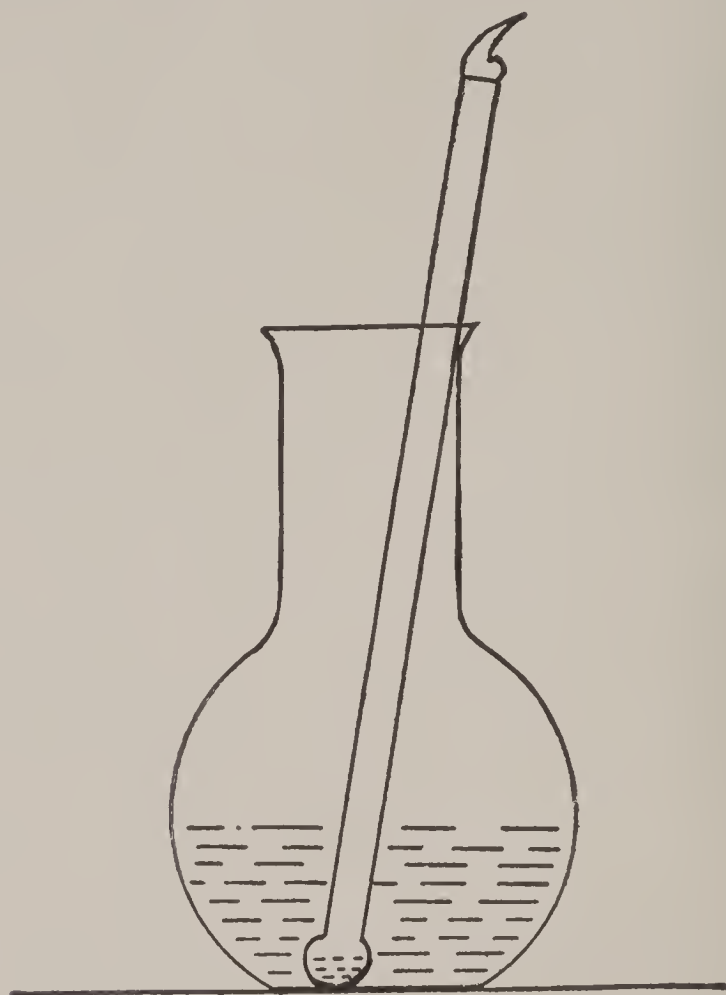


FIGURE 34.

Boiling ether with the heat liberated by the crystallization of sodium sulfate.

this an interesting operation, but you will not succeed at first. Select a piece of glass tubing about a foot long and having as thick walls as possible. Put one end of this in the flame and holding the tube in a nearly vertical position soften the end and let it fuse into a solid mass. When you have formed a solid mass of

molten glass and the opening has been completely closed, remove the tube from the flame and gently blow into it. If you are careful, you can produce an evenly rounded bulb of considerable size. A little practice will make you an expert. When you have produced a good piece of work, fill the bulb with ether or carbon disulfide. You now have an ether thermometer. If placed in a liquid having a temperature of 36° C. or more, the ether will boil.

When the solution of sodium sulfate has thoroughly cooled, immerse in it the bulb of your thermometer and give the flask a quick shake (Fig. 34). The salt will turn to a solid mass of crystals and produce enough heat to boil the ether and cause it to rise in the tube. If you have difficulty in crystallizing the salt by shaking, add a tiny crystal of the salt itself. This gives a center about which crystallization may start and will do the trick. If you like, you may ignite the ether vapor at the top of the tube.

CHAPTER 33

GUGLIELMO MARCONI

WHEN Marconi, an Italian youth born in 1874, was twelve years old, Hertz demonstrated the possibility of signaling through space by means of ether waves. This discovery created a tremendous interest throughout the whole scientific world. But it remained for Marconi, while still little more than a boy, to make the first revolutionizing application of these waves and demonstrate the wonderful possibilities of wireless communication. Patiently he waited for the learned scientists in their well equipped laboratories to invent and perfect wireless apparatus, but he waited in vain. At length he started to experiment for himself, and soon he startled the world.

In his father's garden he set up aërials and crude instruments of his own devising. Then making one discovery after another he began actually to send and receive signals. At first he could send over distances of only a few hundred feet. But he worked over his apparatus by day and dreamed of it by night. Such enthusiasm and earnest effort were bound to produce results, and they did. By 1896 he was sending over distances of several miles. Then he went to England where he made remarkable demonstrations. He estab-

lished a station on the Isle of Wight and maintained communication with the mainland. He sent messages from ship to shore and in 1898 reported for a Dublin paper the news of the annual Kingstown regatta. He established wireless communication across the Channel between England and France.

Then Marconi won his most notable triumph. He came to America in 1901. On the coast of Newfoundland in the strong winds and bitter cold of December he raised a huge kite-aerial high in the air and led his lead-in wire to simple instruments on a table in an upper room of an old barracks. Patiently he waited the signal of the letter "S" to be sent from his station in Cornwall, England. And presently it came. Unmistakably he heard in his telephone receivers the three dots of the pre-arranged signal. Transatlantic wireless was an established fact, and Marconi, then but twenty-six years old, had made a name for himself which will never perish. The vision of the boy experimenter had been translated into a wonderful achievement. And Marconi is still experimenting.

CHAPTER 34

SOME EXPERIMENTS IN PHYSICS

Air Pressure.—Obtain a gallon can and fit the neck with a snug cork. Remove the cork and pour into the can enough water to cover the bottom well. Place the can over the Bunsen burner and bring the water to a boil. When the steam is issuing vigorously, insert the cork and quickly thrust the can under a cold water faucet or into a pail of cold water. The steam will condense on the inside, leaving a vacuum, and the pressure of the outside air will crush the can into an unrecognizable mass of tin.

Floating Copper on Water.—Despite the fact that copper is nearly 9 times as heavy as water, you can float a copper wire. Allow a basin of cold water to stand for a short time. Cut several lengths of fine copper wire. Holding a piece two or three inches above the water drop it. The wire will curve the surface of the water downward and float in a shallow depression. Owing to what we call surface tension the surface of a body of still water behaves as though a tight membrane were stretched over it.

Singing Flames.—Obtain two three-quarter inch gas pipes about 18 inches long. Clamp each in a vertical position to a ring-stand and place it over a Bunsen

burner. Light one of the burners and lower the pipe over the flame. Then regulate the gas pressure until you get a musical note. In the same way start the other flame to singing. Make a paper cylinder about 8 inches long and just large enough to slip over the pipes. With the two flames singing raise and lower the cylinder over one of the pipes until you will hear a distinct throbbing. This alternate increase and decrease of the sound is shown as beats. By changing the length of the pipe you change the pitch of the sound produced, and, when two notes of different pitches are sounded together beats are heard. This is one of the chief causes of discord in music.

If you have trouble in getting the Bunsen flame to sing, soften a piece of glass tubing in the flame and draw it out so as to make a jet tube. Substitute this for the burner, connecting it to the gas supply and lowering the pipe over the flame.

Musical Pipes.—Secure a lamp chimney about 14 inches long having straight sides and as small a diameter as possible. Clamp this in a vertical position to a ring-stand. Find a short length of thin walled brass tubing and flatten one end so that only a narrow slit is left. Connect the opposite end to an 18-inch length of rubber tubing.

Blow across the mouth of the mounted chimney. A musical note will be obtained. You will need a little practice in order to know just how hard to blow and how to direct the air jet. Rest the chimney on the

table top and blow again. Now you have a closed pipe. You will notice that the tone is much lower in pitch. It is just an octave lower.

Now fill a wide mouth bottle nearly full of water and lower the chimney into it. Adjust the chimney until the length of the air column (from the top of the chim-

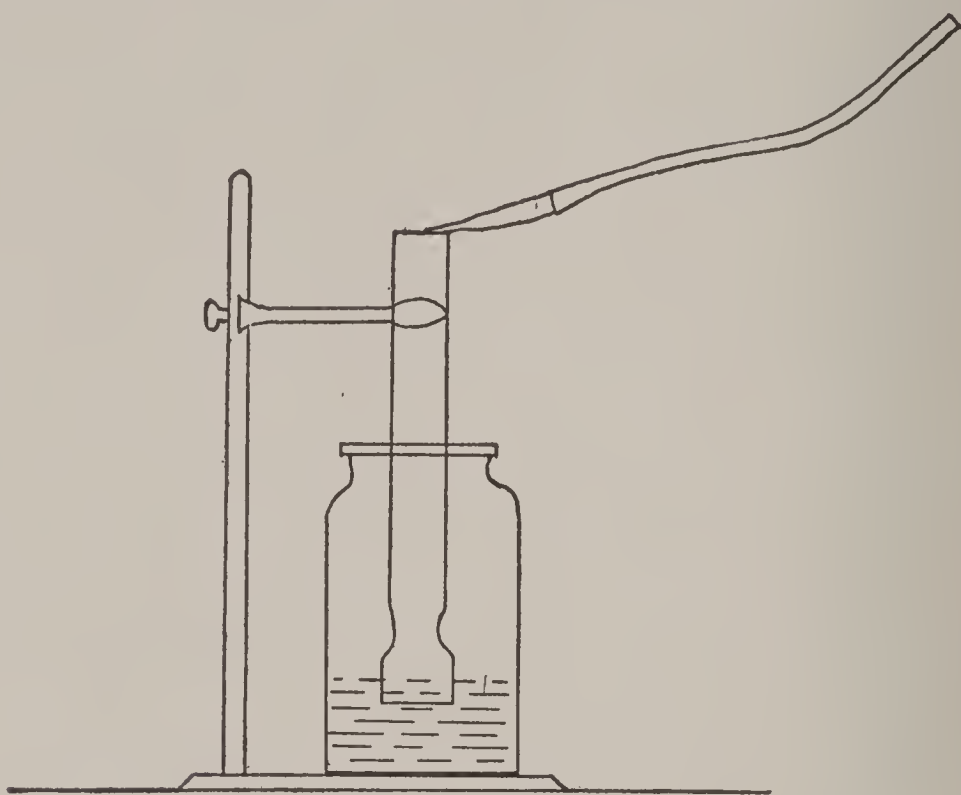


FIGURE 35.

Producing a musical note by blowing across the mouth of a closed pipe.

ney to the water) is 13.25 inches. Blow across the mouth and you will obtain the note known as middle C (Fig. 35). Lower the chimney until the air column is 11.75 inches long and blow again. Now you obtain the next note of the scale. Continue to adjust the chimney, making in succession lengths of 10.5, 9.9, 8.8, 7.9, 7, and 6.6 inches respectively. These various

lengths beginning with 13.25 and ending with 6.6 will give you the notes of the scale.

If you can arrange a row of bottles, test tubes, or chimneys so as to secure all of these lengths at once, you may play the scale upon them. If you understand music you may play familiar airs. Such a stunt forms a most interesting and novel feature for an evening's entertainment.

An Air Thermometer.—Blow a bulb on the end of a glass tube just as you did in making the ether thermometer in the chapter on crystals. As before, select a piece of tubing having as thick walls as possible. If the first blow does not give a good bulb, remelt the glass and try again. Get as big a bulb as you can, but do not let the walls get too thin. Thrust the tube through a 2-hole stopper fitted into a small bottle or test tube containing colored water. Have the tube dip beneath the surface of the water (Fig. 36). Brush a Bunsen flame across the bulb once or twice. This will expand the air and force some of it out. When the bulb cools water will be drawn up the tube. Then as the temperature changes the water will rise and fall in the tube. As the air in the bulb cools and contracts, the water will rise. As it warms and expands, it will fall. Put a few drops of ether, carbon disulfide, or gasoline on the bulb and watch the water rise. The liquid cools the bulb by evaporation. Put a few drops on your hand and note the cooling effect.

“Burning” Water.—Pour a little ether or gasoline on the surface of water in a beaker or evaporating dish.

Remove the ether or gasoline bottle to a safe distance. Then light the liquid. Apparently the water will take fire and burn.

Boiling water without a flame is an interesting experiment. Fill a 500 c.c. flask half full of water and



FIGURE 36.

Type of air thermometer invented by Galileo.

bring it to a boil, driving out all the air. When the steam is coming well, turn off the gas, stopper the flask with a solid cork, and quickly invert it, supporting it by a tripod or ring-stand set in a large basin. Now pour cold water on the flask and the water will boil vigorously (Fig. 37). The water on the outside of the flask condenses the steam inside, making a vacuum

and causing the water to boil at less than the usual temperature. If the flask is not a strong one, it will sometimes break.

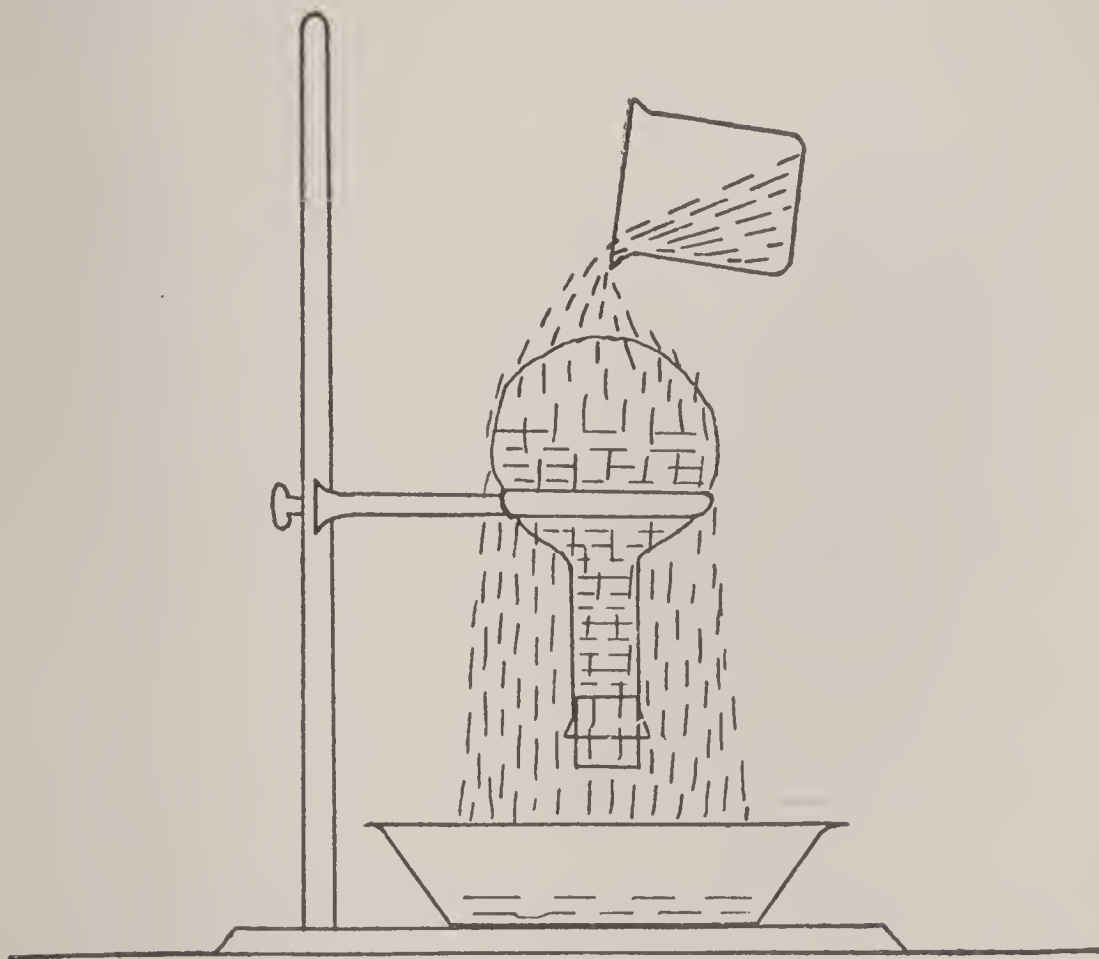


FIGURE 37.

Boiling water at temperatures considerably below 100°C .

The production of heat by chemical action is seen in the mixing of sulfuric acid and water. Pour concentrated sulfuric acid in a thin stream into a test tube half full of water. Note that the test tube becomes unbearably hot.

Heat from the electric current may be well shown, if you have a lighting circuit in your laboratory. Stretch 12 feet of No. 28 iron wire from one

binding post to a wire suspended from the ceiling or other support and back to the other binding post. Darken the room and turn on the current. The wire will become red hot and sag just as telegraph wires do in the summer. Turn off the current and the wire will cool and shorten.

Extremes of temperature in the same vessel may be shown with a test tube of water, a piece of ice, and a lump of lead or a coil of wire. Drop a piece of ice into a test tube and upon it the lead or push down the tube just above the ice a tightly fitting coil of wire. Fill the tube with water. Hold the upper portion of the tube in the flame. The water will boil there, but in the middle portion it will be cold, and in the bottom there will be ice. This proves that water is a poor conductor of heat.

Conductivity of Metal, Glass, and Air.—Obtain two pint cans and a glass beaker of equal capacity. Set one of the cans into a small box containing a half inch layer of cotton wool or boiler felt. About the sides of the can stuff the box full of the packing material. Leave the other can in the open. Fill each can and the beaker with boiling water and cover each with a glass plate or square of asbestos. At intervals take the temperature of the water in each container. In which does it cool fastest? Slowest? The principle of the fireless cooker is illustrated in the can packed in cotton wool. Both the cotton and the wooden box are heat insulators, but the best insulator is the air in the spaces between the fibers of cotton.

Convection Currents.—Heat is distributed through liquids and gases by currents set up by differences in density due to differences in temperature. Pour a quarter-inch depth of water into a tumbler and set in it

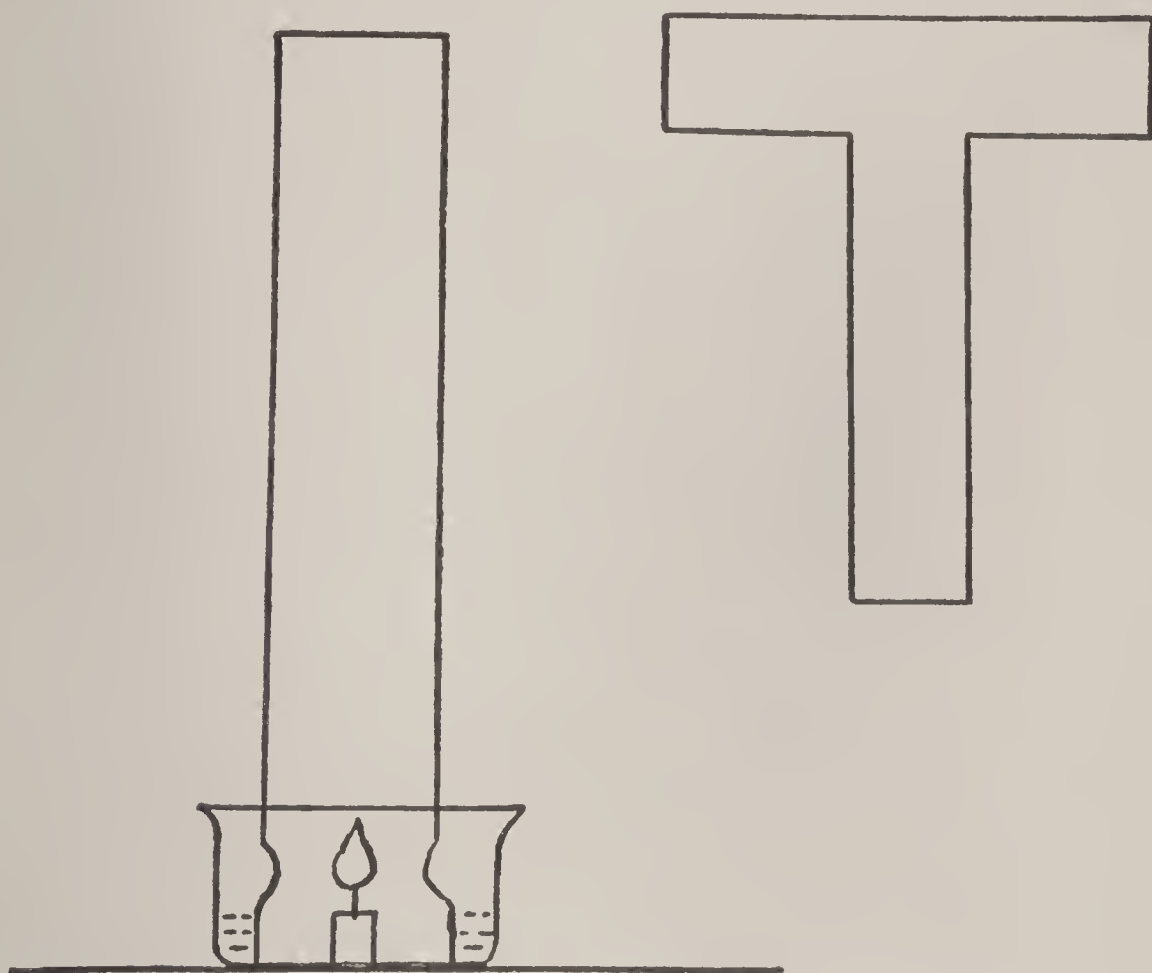


FIGURE 38.

Apparatus for showing convection currents.

a candle about an inch long. Light the candle and place over it a lamp chimney (Fig. 38). In a moment or two the flame will go out. It has burned out the oxygen and produced an atmosphere of carbon dioxide. Again light the candle and place over it the chimney, but divide the chimney by a cardboard partition ex-

tending about two-thirds the way down its length. This time the flame will not go out. It will burn indefinitely and by its fluttering you will see that a current of air is passing through the apparatus. Hold a lighted match or a piece of touch paper on each side of the partition and you will see that air is passing down one side of the chimney and rising on the other. This experiment illustrates the principle of hot air heating. The air at the bottom of the chimney expands with the heat and becomes lighter than the cooler air above. Therefore the cold heavy air settles on one side of the partition and pushes up the warmer lighter air on the other side.

Freezing Mixtures.—In a metal cup or can mix 1 part of salt with 3 parts of crushed ice or snow. Set the can on a block of wood having on the top of it a little water. Very shortly you will observe moisture collecting on the outside of the can and then a deep layer of frost. A thermometer thrust into the mixture will show temperatures considerably below zero. Pick up the can and you will find it frozen very securely to the block. Water in a test tube thrust into the ice will quickly freeze.

By mixing 3 parts of calcium chloride with 2 parts of snow or ice, you may obtain a temperature low enough to freeze mercury. This means 40 degrees below zero.

Making a Lifting Magnet.—An electromagnet that will prove to be a little giant and one which you may connect directly to a 110-volt circuit may be easily



Lifting 45 pounds with a homemade electromagnet. Its capacity is more than 200 pounds.

made. Go to a blacksmith shop and get a rod of perfectly *soft iron* about 18 inches long and 1 inch in diameter. Have the blacksmith bend this into the shape of the letter U. From heavy leather cut 4 washers each about 5 inches in diameter and just large enough to fit snugly on the iron rod. Thrust two of them well up on the arms of the rod and place the others nearly at the bottom. Wind each arm between these washers with tape. Then buy several pounds of No. 26 double covered copper wire. Leaving long lead wires at the end, wind this wire as tightly and closely as possible between the washers on one arm. When this side has been filled carry the wire across and fill the space between the washers on the other arm. But in doing so be very sure that you wind the wire on the second arm in the same direction that you did on the first. That is, wind it so that if the rod were straightened out the wire would be wound in the same direction from end to end. To wind this wire will require considerable time, but the more you wind, the greater will be the lifting power of your magnet.

When the wire is wound, make a soft iron hanger to fit across the poles. This should be a strip of quarter inch iron 1 inch wide and carrying a hook at the center.

If you have used sufficient wire, the magnet may be safely connected to the 110-volt lighting circuit. But, if in doing so, you notice that the winding seems to get hot break the circuit at once and, before using again, connect in series with it your lampboard rheostat, and turn on enough lamps to give you the maximum lifting

power without heating. Try its lifting power on bars of iron, nails, and spikes. If you have inserted a switch in the circuit, you may at any moment drop the load by opening the switch. The fact that an electromagnet may be made instantly to acquire or lose its magnetism is one of its chief advantages. Hang the magnet on a crowbar placed across two saw horses. Put the hanger across the poles and close the circuit. Attach to the hook the heaviest load you can find. You should be able to lift at least 100 pounds and probably more.

A Pinhole Camera.—For this experiment you must have a room which you can make perfectly dark, and the shades at the windows should be black so that no light will come through them. Blankets may be hung over all but one of the windows, if necessary. Through the dark shade of this window make a hole about an eighth of an inch across and hold an 18-inch square of white bristol board 1 foot from it. If there is good sunlight outside, you will obtain an excellent image of the landscape and whatever happens to be within range. But you will notice that the image is inverted. This is always true of what we call *real images*. The rays of light reflected from the distant objects pass in straight lines through the small opening and come to focus on your screen. But in passing through the opening in the shutter, the rays cross and this inverts the image.

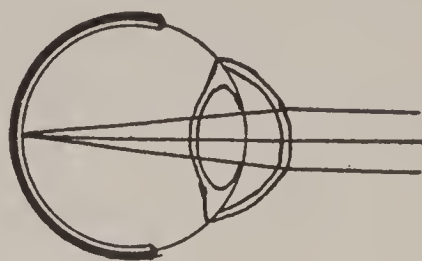
“Burning” a Candle in a Beaker of Water.—Clamp in a vertical position on your “magic demonstration table” a pane of window glass. Raised on a block of

wood about a foot in front of the glass set a candle. On a block of the same height and at the same distance behind the glass set a beaker, or bottle, of water. Light the candle and by reflection it will seem to be burning in the beaker of water. Light from the beaker is transmitted through the glass to your eye, and light from the candle is reflected from the glass to your eye. This image is upright as you see, and is called a *virtual image*. It is the same size as the object and as far behind the mirror as the object is in front. Move to either side of the room and find how far you can go and still see the image of the candle. The wider the pane of glass, the greater this distance will be.

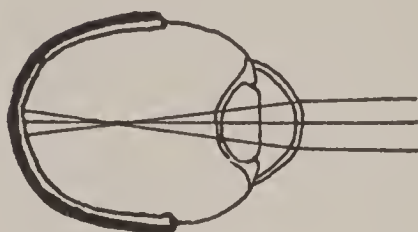
Near-sightedness and Far-sightedness.—For this experiment you will need a double concave lens of about 15 centimeters focal length and two double convex lenses of about 20 to 25 centimeters focal length. Make a wire holder for one of the convex lenses and mount it on a yardstick. Standing a little ways back from an open window, hold the lens so as to catch the light coming through the window and place behind it a white cardboard screen. Move the screen until you obtain a distinct image of the landscape. The distance between the lens and the screen will be the focal length of the lens. The crystalline lens of the eye produces an inverted real image upon the retina in exactly the same way.

Now in a darkened room place the yardstick lengthwise of the table and secure it in position with blocking at one end. Place a lighted candle on one side of the

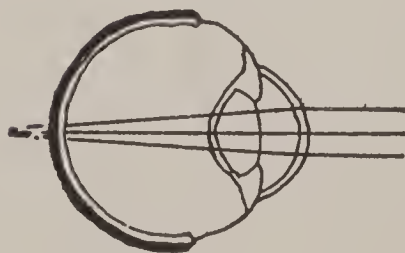
lens at a distance of about twice the focal length. On the other side place the screen and move it back and



A



B



C

FIGURE 39.

A shows a normal eye with the image of a distant object brought to focus on the retina; B shows the condition for a near-sighted eye; and C shows it for a far-sighted eye.

forth until you obtain a clear image. The lens corresponds to the crystalline lens of the eye and the screen to the retina. Unless the image on our retinas is clear

and distinct, our sense of vision will be blurred. Move the screen toward the lens. The image is now indistinct. Move it a little back of its original position and the image will again be indistinct.

The lens of the eye is the most remarkable lens in the world. In a normal eye it is able to change its focal length according to the distance of the object so as always to bring the image to focus on the retina. But in a near-sighted eye, unless an object is held closer to the eye than is necessary for a normal individual, the image will come to focus in front of the retina. The result is a blurred impression. In a far-sighted eye, the image comes to focus behind the retina, unless the object is held farther away from the eye than is necessary for normal sight. You illustrated near-sightedness when you held the screen farther away from the lens than the position for a distant image, and far-sightedness when you held it nearer to the lens.

To show how near-sightedness is corrected, hold the screen a little behind the position necessary for a clear image. Then hold between the candle and the lens the double concave lens. Move it back and forth until you obtain a clear image on the screen. That is what glasses do for near-sighted people. Concave lenses diverge the rays of light just enough to bring the image to focus on the retina instead of in front of it.

To show how far-sightedness is corrected, hold the screen a little in front of the position necessary to se-

cure a distinct image. The image will be blurred. Then hold between the candle and the lens the other convex lens. Move it back and forth until you obtain a clear image on the screen. The convex lens converges the rays of light and brings them to focus on the retina.

A Compound Microscope.—A compound microscope consists of an objective and an eyepiece. Both are convex lenses. The objective is of shorter focal length than the eyepiece. To make a microscope select two double convex lenses, one of about 5 cm. focal length and the other of about 10 cm. You can always determine the focal length by focusing the image of a distant object on a screen and measuring the distance between the image and the lens. Mount the objective in the jaws of a clamp attached to a ring-stand and place it just a little more than its focal length from the table top. In another clamp directly above the first mount the eyepiece so that the centers of the lenses are in the same straight line. Now place a sample of fine type on the table beneath the objective and move the eyepiece up and down until you get a distinct image. The farther you can separate the lenses and still get a clear image, the greater will be the magnifying power of your microscope. You will observe that the image is inverted. If you wish to move the image in one direction, you must move the object in the opposite direction.

Scientists have already reached the limit of magnification with compound microscopes. The limit for

practical work is about 1200 diameters, although magnifications of 3000 diameters have been obtained.

A Galilean Telescope.—Galileo, a famous Italian scientist, first explored the heavens with a telescope in 1608. With his simple instrument he made marvelous discoveries. He revealed the moons of Jupiter, he disclosed new stars, and resolved the beautiful Milky Way into a countless host of infinitely distant suns.

The common opera glass is made on the same principle as Galileo's telescope. It will be easy for you to arrange lenses to illustrate this telescope yourself. Mount in a wire holder on a yardstick a double convex lens of about 10 to 15 inches focal length for the objective and a double concave lens of 4 to 6 inches focal length for the eyepiece. Adjust the lenses so their centers are in a straight line and point them toward some not too distant object. Move the eyepiece back and forth until you get a distinct image. You will observe that the image is upright and that you secure considerable magnification, or better distant objects seem to be brought nearer.

An Astronomical Telescope.—It is entirely possible for any boy to make an astronomical telescope of considerable power and one which will afford him a great deal of pleasure. The first consideration in making a telescope is to secure the lenses. You will need two—a double convex lens of long focal length for the objective and one of short focal length for the eyepiece. The greater the difference in these focal lengths, the greater will be the magnifying power. If the objec-

tive has a focal length of 36 inches and the eyepiece is 2 inches in focal length, the telescope will magnify 18 times. And those numbers will be about the proper lengths. The eyepiece may be from 2 to 4 inches.

The next essential is a heavy cardboard cylinder just large enough in diameter to take the objective and as long as its focal length. At a stationer's or some supply house dealing in wrapping paper and mailing devices possibly you can find such a cylinder. If not, you will have to make a cylinder. To do so cut a length from a curtain pole or some other round stick of wood having the same diameter as the lens. Then obtain a number of sheets of heavy wrapping paper and prepare a pot of glue or paste such as a paper-hanger uses. Cut the paper to the exact length of the pole. Lay a strip of it on the table and roll it once about the pole. With a brush spread paste on the unrolled portion of the strip and roll it very tightly about the pole. Repeat this with other strips until you have a cylinder of the proper thickness. After it has dried well, remove the pole and your telescope tube is ready.

You must now mount the objective. Just inside one end of the tube paste a narrow ring of cardboard, being sure that the distance between it and the end of the tube is everywhere the same. Slip the objective into place letting it rest against this cardboard ring. Secure it there with another cardboard ring.

Now you are ready for the eyepiece. Since this lens is of smaller diameter, it must be mounted in a smaller

tube. Prepare another paper cylinder just large enough to hold the lens and about 3 or 4 inches longer than its focal length. Place the eyepiece in one end just as you did the objective, but to enable you to see to the best advantage only the center portion of the glass should be exposed. Therefore, directly in front of the eyepiece place a cardboard disk having a small hole about an eighth to a quarter of an inch in diameter exactly in its center. Just inside the big tube place a disk of wood about an inch thick and having a hole bored in it just large enough to take the smaller tube and to permit of its being freely moved back and forth.



FIGURE 40.

Astronomical Telescope.

Insert the eyepiece tube and your telescope will be ready for its first try-out. (Fig. 40.)

For distinct vision the eyepiece should be drawn out until the distance between it and the objective is equal to the sum of the focal lengths of the lenses. Point the instrument toward a distant object and bring it into focus. Of course it will be magnified and appear much nearer, but you will be surprised possibly at one thing. The image will be inverted. In an astronomical telescope this makes no difference. It does not matter if the moon is upside down. In field glasses, however, another lens system must be used to

reinvert the image. This results in some loss of light and is unnecessary in an astronomical telescope.

You will find it difficult to hold your instrument steady enough to keep an object in the field of vision. You will need a support. Three small curtain poles or old broom sticks and a little ingenuity will enable you to contrive one.

On a clear night explore the moon. It is the most beautiful object of the heavens. Examine its craters and rugged mountains. I am sure you will be delighted with the results. Of course you cannot magnify the stars. Their distances from us are too immense. But the planets Venus and Jupiter will be well worth your attention.

CHAPTER 35

"THE HALL OF FAME"

DURING my college days the members of my fraternity carried out a novel stunt in the way of evening entertainment. In the attic of the "frat" house we made a plaster of Paris cast of each member's face. Then on the walls of a small room draped with black cloth and dimly lighted with candles, so as to produce as weird an effect as possible, we hung the casts. Each one was numbered, and each guest was required to pass judgment upon the identity of each cast. In a little booklet he recorded his guesses, for in many instances they were little more than that. It is no easy matter correctly to connect a particular cast with the owner's face. If you did not know your own, you quite likely would be at a loss to pick it out.

This performance is really of a chemical nature, for we must know that plaster of Paris is gypsum with half of its water of crystallization driven off, and that when it sets it unites chemically with this water again.

To make these casts, select some old shop or room in which the muss which it will make will not matter. Have at hand a large quantity of plaster of Paris, a supply of water, and plenty of vaseline. Each aspirant to the honor of being included in this "hall of

fame" must first grease his face *thoroughly* with vaseline. Be sure that the eye-lashes, eyebrows, and every part that will be touched by the mixture are well covered. If not, the experience in removing the first impression will be very painful. I have seen boys almost weep at this part of the operation.

An old bench or cot may serve as operating table. Lay the victim upon it, wrap a towel about his face to hold the plastic mass in place, and do not forget to insert in his nostrils two paper tubes, through which he may breathe while cut off from the outer world. Apply a rather thick mixture of plaster of Paris and water liberally, making a layer at least an inch thick. For the patient there will be perfect darkness, and he will be unable to speak. Presently he will have evidence that a chemical change is taking place in the mixture on his face. He will feel the gentle warmth from the heat developed in the chemical union of the water and plaster of Paris. After about 15 minutes, the mixture will have set, and the cast may be removed. To do this stand up, hold your hands over the cast, and shake and pull it off. If the face has been well vaselined, this will be a very simple matter.

You now have the form in which the finished cast will be molded. Vaseline the inside of this mold thoroughly and fill it with another mixture of the plaster of Paris. While the mixture is still soft insert a wire loop with which to hang the cast in the "hall of fame." When this second mixture has hardened, break away the first cast by pounding it gently with a wooden

mallet. Retouch the cast by filling all the pores and cavities with some fresh mixture and paint it with white enamel. You will be delighted with the result.

Hang these in the chosen room. If it is properly draped and lighted, the effect may be made weird and ghostlike. I have never seen a more unusual or more interest provoking event at an evening's entertainment than this. Try it.

SOURCES OF SUPPLY

Some of the apparatus and material needed to perform the experiments described in the foregoing pages may be found in the home or had from the grocer and the local druggist. Real apparatus and many of the chemicals must be purchased from some supply house. Following is a list of firms located in various parts of the country:

Cambridge Botanical Supply Co., Waverly, Mass.

Central Scientific Co., Chicago, Ill.

W. M. Welch Scientific Co., Chicago, Ill.

Denver Fire Clay Co., Denver, Colo.

Eimer and Amend, New York, N. Y.

Braun-Knecht-Heimann Co., San Francisco, Calif.

Possibly some of these companies will send you catalogues, if you request it and explain that you are starting a home laboratory. Select the experiments which you wish to do first. Then purchase your equipment. Let it be good and take care of it.

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